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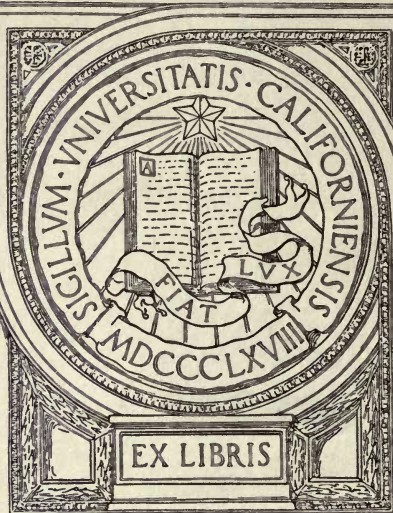


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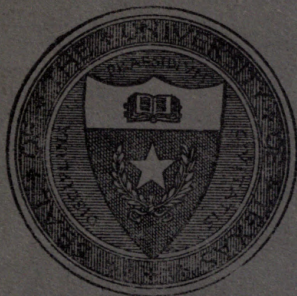
OFFICIAL SERIES No. 64

DECEMBER 8, 1911

Chemistry in High Schools

BY

E. P. SCHOCH, Ph. D.,
Professor of Physical Chemistry
The University of Texas



PUBLISHED BY
THE UNIVERSITY OF TEXAS
AUSTIN, TEXAS

Entered as second-class mail matter at the postoffice at Austin, Texas

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Acknowledgment.

Many teachers of science in Texas schools have kindly favored the writer with expressions of their experience on several topics considered in this paper, and their collective advice has been carefully incorporated. Indebtedness is hereby gratefully acknowledged.

WHAT SOME OF OUR SMALLER TOWNS HAVE DONE FOR
CHEMISTRY.



Chemical Laboratory, Winnsboro (Texas) High School.

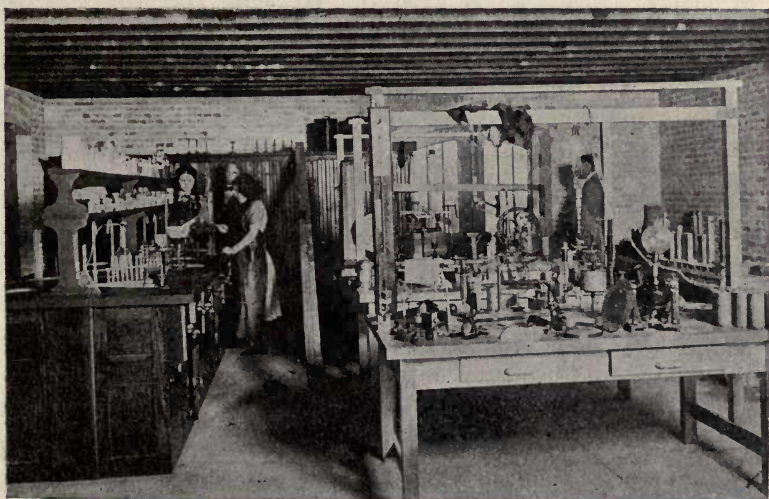


Chemical Laboratory, Corpus Christi (Texas) High School.

WHAT SOME OF OUR SMALLER TOWNS HAVE DONE FOR
CHEMISTRY.



Chemical Laboratory, San Marcos (Texas) High School.



Chemical Laboratory (on left), Eagle Lake (Texas) High School.

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CHEMISTRY IN HIGH SCHOOLS.

PART I:—EQUIPMENT.

The Teacher

First and most important is the teacher. He should be above all a *science* man, both by inclination and training,—a person whose inclinations and desires naturally make him fond of the subject. He should have had some sound training in chemistry and he should have received this training at a place where that subject is taught *extensively* rather than at a place where only introductory work is given, since much necessary information is absorbed from the proper “atmosphere.” The extent of the person’s training in chemistry should not be less than two thorough courses in a first-class college or university, and in addition some college training (at least one course) in either physics or in a biological subject (botany or zoology). This minimum requirement is not excessive, and frequently persons may be obtained who have had much more training than this.

When and Where to Order Laboratory Supplies.

The teacher who is to use the laboratory should design its equipment and order the materials.

The supplies should be ordered early in June. It is quite profitable to submit large orders to several of the large importing houses for quotations for *duty free* importation and *delivered at your railway station*. If the ordering is delayed beyond July 1st it will scarcely be possible to secure a duty free importation by the time the session opens in the fall. All of the large chemical dealers mentioned in the list below will import apparatus duty free if requested to do so.

C. H. Stoelting Co., successors to Chicago Laboratory

Supply & Scale Co., 31-45 W. Randolph St. Chicago, Ill.

Central Scientific Co., 20-28 Michigan St.....	Chicago, Ill.
Edward P. Martin Co., 144-146 E. Erie St.....	Chicago, Ill.
L. E. Knott Apparatus Co.....	Boston, Mass.
Bausch & Lomb Optical Co.....	Rochester, N Y.
Woldenberg & Schaar, 1025 S. State St.....	Chicago, Ill.
Eimer & Amend, 205-211 3rd Ave.....	New York.
The Kny-Scheerer Co., 404 West 27th St.....	New York.
E. H. Sargent & Co.....	Chicago, Ill.

Selection of a Room for the Laboratory.

In the selection of a room for the chemical laboratory the following requirements must be met:

The room must have such means of ventilation that all fumes may be removed rapidly and not blown into other parts of the building.

Communicating with the laboratory there must be a small room where the supplies may be kept under lock, and in which the teacher may prepare experiments and solutions.

The store and preparation room in turn must communicate with the class room so that the setting up and removal of lecture apparatus may be facilitated.

If the teacher also teaches physics, then the store and preparation room, as well as the class room, may be used for both subjects; but in that case the physical laboratory should be located conveniently near so that apparatus may be transferred to and from the store room without unnecessary trouble.

The simplest way to meet the requirement of an adequate draft is to select a room with outer walls and windows on at least two sides. If an inner room with only one exposure (south!) must be used, then there must be provided at the farther end of the room from the windows one or more ventilating shafts so located that all parts of the room may be swept by the air current coming in at the windows and passing out through the shafts. Such a shaft should be one to two feet in diameter, and it should extend straight to the top of the building. The cover or cap on the roof over the top of this shaft should be placed very high (say twelve inches) above the top of the shaft, so that it will not impede the draft. Cornice men frequently

place this cap entirely too low. It is desirable to aid the draft by putting a gas flame or an electric fan at the bottom inlet.

The form of ventilators frequently employed for ventilating ordinary buildings (that is, narrow shafts extending up in the wall, and opening into the rooms by means of a series of holes near the ceiling of the room) provides *insufficient ventilation for a chemical laboratory*.

Draft Hoods.

Obnoxious gases should not be allowed to escape freely into the atmosphere of the laboratory, but should be confined to special draft hoods. Such draft hoods must be carefully designed, otherwise they are worse than useless. The following rules should be observed in their design:

1. The hood should be as small as practicable. More and smaller hoods should be the rule.

2. The stack should be vertical throughout its entire length.

3. The hood proper, which is the part below the stack in which the apparatus is placed, should be as small as possible, so that the velocity with which the air moves through it may not be less than one-third of that with which it moves in the stack. The velocity of the draft should be great enough to lift the heaviest vapors, as, for instance, those of sulphuric acid. The inlet for the draft should be placed so that the current of air from that point to the chimney may sweep through all parts of the hood. For this purpose some stops are usually placed so as to *prevent closing entirely* the front door of the hood. This carefully adjusted air inlet is essential to the successful operation of the hood. The hood proper should be made as low as convenience will permit in order that the length of the *slowly moving air column in the hood* may be as short as possible.

4. The stack or chimney should be proportionate to the size of the hood and its cross sections should not be essentially less than one-third of the cross section of the hood proper.

5. A large gas burner should be placed at the base of the stack at a point at which it is possible to light the gas by reaching up into the hood with a burning candle on a stick. This gas flame need not always be lit when the hood is used, but in

Fig. 1a.

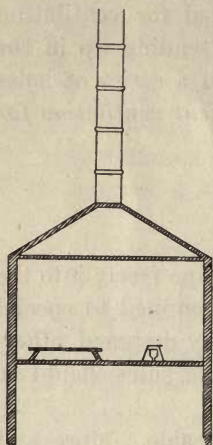
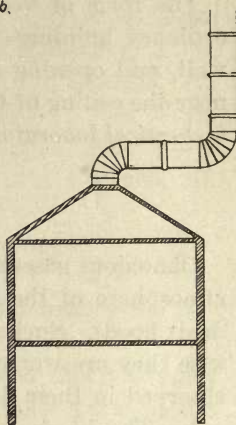


Fig 1 b.



Examples of Faulty Hood Design.

Fig. 2a.

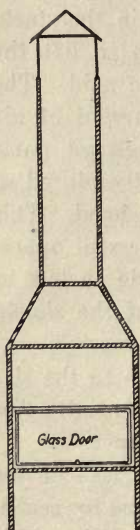
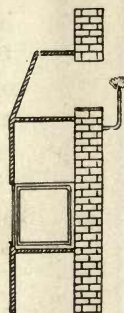


Fig 2b



Front



Side

Examples of Proper Hood Design.
From the *Chemical Engineer*.

quiet weather it must be lit in order to establish a sharp draft in the chimney.

6. A cap covering the stack for the purpose of keeping out rain is frequently not necessary and naturally the hood will operate best without the cap. Should a cap be found necessary, then care must be taken to place it very high above the top edge of the stack in order that the draft may not be impeded by it.

Hoods are frequently constructed without any consideration of the fundamental principles to be observed in order to get a good draft. Fig. 1-a shows a hood in which the stack is entirely too small. In Fig. 1-b, the stack suffers from the further impediment of not being straight. These mistakes in design are frequently found in actual hood construction. Fig. 2-a shows the ideal hood construction. In this hood the stack runs straight up from the closet and is large enough to provide for a sharp draft. Fig. 2-b shows a mode of construction in which the closet is built to connect with a large flue in the wall. If the *diameter of this flue is large enough* and a flame is placed as shown, then the hood will operate properly. All these figures show the form in which the hood proper, or closet, is generally built. The bottom of the closet is usually at the height of the ordinary laboratory desk and the height of the closet above the table to the contracting portion is 30 to 36 inches. The sliding front door, as well as the two sides, should be made of glass.

Arrangement of Desks, Shelves, etc., in the Laboratory.

In the placing of the laboratory desks and the supply shelves care must be taken to leave enough room for the students and instructor to pass each other and secure their supplies readily. The distance between the desks should be 4 ft. 6 in., certainly not less than 4 ft. Desks should not be placed adjacent to the wall anywhere, and a space at least 6 ft. wide should be left between the wall and the desks all around the room. However, this space is large enough to place shelves along the wall wherever desired, as well as shelves for blast lamps, balances, and other special apparatus. It is advisable to place a platform with a suitable table and a black-board in the room so situated that

the instructor may make special demonstrations before the class or give special explanations while laboratory work is in progress.

In placing the desks, the manner of running the gas, water and sewer mains should be considered: It is most desirable to run these *underneath the ceiling of the room below* the laboratory—hence open to view and accessible for repairs at any time; this permits the placing of the desks in any way desired. If, however, the pipe mains must be placed *in the floor*, then they should run so as not to cross the floor joists, and the desks must be placed accordingly. The floors over the pipe mains should not be nailed down again, but made “trap-door” fashion.

There should be *no high shelves* or other super-structure *on top* of the laboratory desks which would prevent the instructor from looking across the room and seeing what is on top of any desk. The sets of reagent bottles are to be placed on small, low, movable shelves, 36 to 40 inches in length, designed to hold one row of bottles on each side, with a partition at the centre not over six inches high. The bottoms of the shelves should be six inches “clear” above the table top, but the total height of the shelf and bottles should not exceed 12 to 14 inches.

Details of Desks.

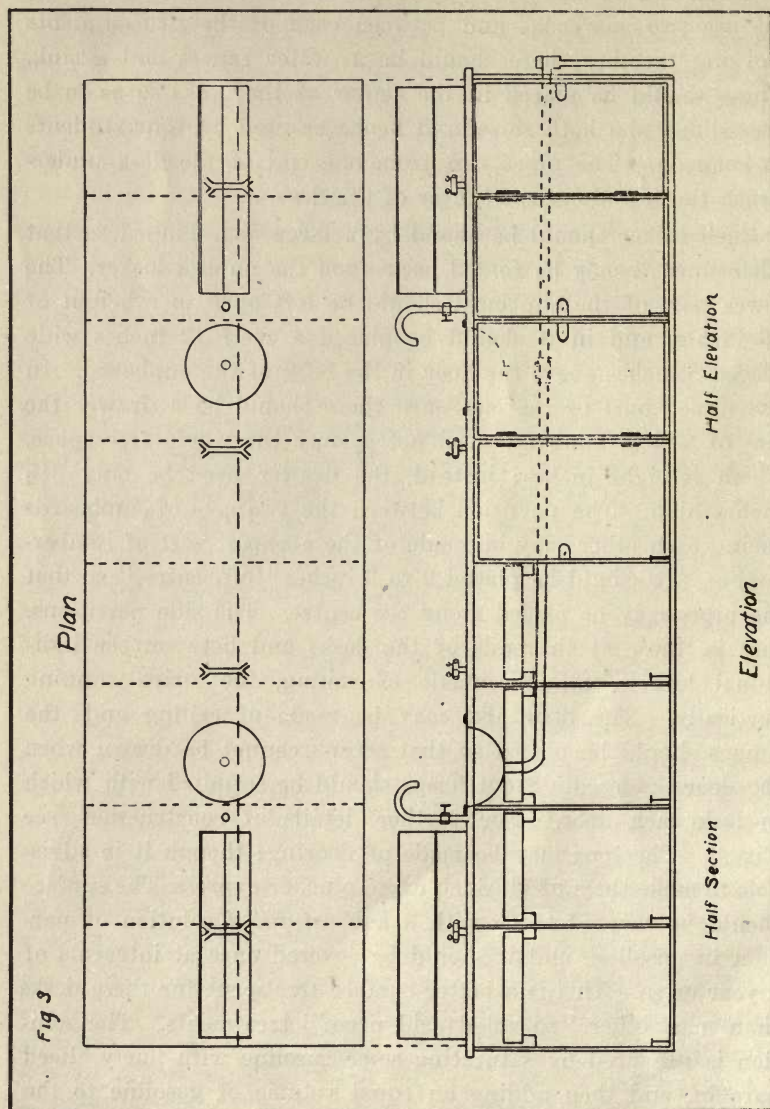
In most schools the laboratory is to be used by two sets of students on different days. This is easily arranged, because the amount of desk top space required by a student while working is large enough to construct two independent lockers underneath it. In some cases instructors have even arranged for three independent lockers under this space so that the laboratory may be used by three distinct groups of students at different periods. They place four deep drawers in the space of the two lockers used in the former arrangement:—three of these are locked separately and in them are placed the perishable apparatus dealt out to the students individually. The fourth is not locked and in it are placed the iron ring stand, the wooden funnel stand, the burners and other large and usually non-breakable apparatus to be used by three students in common. This arrangement is not as convenient as the *two student* arrangement.

The table top space allowed per student, while at work,

should be 42 inches in width, and 24 inches in depth. The desks should be 36 inches high. Each student should have for his use two gas cocks and between each of the two students working together there should be a water faucet and a sink. These should be placed in the center of the desks so as to be accessible from both sides, and hence be used by four students in common. The pipes run from one end of the desk underneath the top along the center of the desk.

Each locker should be closed by a large door hinged so that when open it may be folded back upon the unused locker. The lower part of the cupboard should be left open to a height of 26 inches, and in it should be placed a shelf 12 inches wide placed 8 inches above the floor in the back of the cupboard. In the upper part of the cupboard there should be a drawer the size of which, however, need not occupy the whole free space, which is eight inches; instead, the drawer need be only $3\frac{1}{2}$ inches high. The partition between the two sets of cupboards facing each other may be made of the cheapest sort of lumber-boxing. It should be placed 2 to 3 inches "off-centre," so that the pipes may be placed along the centre. The side partitions, that is, those at the ends of the desks and between the individual lockers, may be made of ceiling, the strips running vertically. The door also may be made of ceiling and the hinges should be placed so that screws cannot be drawn when the door is closed. Stout hasps should be supplied with which to lock each door. For further details of construction see Fig. 3. The tops may be made of flooring; though it is advisable to make them of $1\frac{1}{4}$ inch white pine, or cypress. The surface should be covered twice with a *half-saturated* solution of paraffin in gasoline, and it should be covered once at intervals of a year or so. This is a better surface treatment for these desks than most other "so-called acid proof" treatments. The solution is prepared by saturating some gasoline with finely sliced paraffin, and then adding an equal volume of gasoline to the solution.

The desks should not be too long: 14 ft. is the *maximum* length desirable; such a desk contains 16 lockers, and top



Chemical Laboratory Desk.

space for eight students working simultaneously four on each side. It is advisable to make the desks only half as long, that is, with only 4 lockers to the side, and then to place sinks at the ends of the desks.

Plumbing.

It is unnecessary to "trap" each one of the sinks separately; the main pipe only needs to be trapped at some convenient point; or better still, it may empty into a "hopper." The latter acts also as a sieve to catch solids, etc., and thus prevents the choking up of the sewer beyond the hopper.

The desks should be so arranged that the main pipes pass up the center or one side of the laboratory so as to touch the end of each desk. Of course, the size of these mains depends on the number of desks. In the laboratory to which the accompanying figures refer, built to accommodate 236 individual lockers, the sewer main is a cast iron pipe 4 inches in diameter, the water main is a 2 inch pipe and the gas main is a 2 inch pipe. The laterals for a 16 locker desk unit as used in this laboratory are of the following dimensions: drain pipe, 2 inches (cast iron); water and gas, one inch each. Where the drain pipe comes out of the desks and turns downward, a "T" should be placed instead of an elbow. By this means an obstruction in the pipe may be readily dislodged. Here the pipes should be joined by a *union* in order that a desk may be easily disconnected.

The sinks may be of porcelain: a 14 inch hemispherical bowl such as may be obtained for about one dollar, or a little above, will serve the purpose very well. It should not be fastened to the drain pipe, but the spout should merely extend into a short piece of lead pipe connecting with the drain pipe. Sinks placed at the ends of desks should be of rectangular form, lead lined, 16 by 20, and 10 in. deep inside. A water faucet of the ordinary *bar-fixture* type, will be found suitable. When in use a short piece of rubber tubing should be attached to the faucet to prevent any splashing of water.

The gas cocks should be chosen carefully so as to fit the size of rubber tubing used on the Bunsen burners. The gas cocks commonly used by plumbers for dwellings are *entirely too large* for this purpose.

Cost of Desks and Plumbing.

At present the cost of desks made according to this design is, in Austin, approximately four dollars per locker. The plumbing, including mains, etc., is about \$1.50 per locker additional.

Other Furniture.

One or several hoods will be required. It is difficult to give an estimate of these because the cost depends entirely upon the design. However, the hood proper without the stack need not cost more than \$15 to \$18.

The store-room and the teacher's work room should be supplied with large, commodious, plain shelves, for the stock of supplies, etc. Placed in a separate, well-closed small room, as they should be, they need no doors to keep out the dust, and such open cupboards are much more practical than closed cupboards. The teacher's room should also have a commodious table supplied with gas, and a large sink.

Other Desk Designs and Ready-Made Desks.

In many schools individual desks are now used, and they are arranged so that all the students face in the same direction. Furthermore, a space is left beneath them so that a student can sit conveniently upon a stool while doing his work. Such a desk can be readily designed by taking the space occupied by two cupboards in the former design, adding to it about 20 inches for the open space beneath where the student puts his feet in sitting, thus making a unit of 62 inches width, about 30 inches depth and 36 inches height. The two cupboards are placed on the two sides with the central open space between them. The water, gas and sink may be placed anywhere on the back part of the desk, though they are usually placed in the space between the two lockers and as near the back edge of the top as is possible. The cost of this construction is necessarily greater than that mentioned above. Inclusive of plumbing, it cannot be less than \$20.00 per desk with two lockers, even if it is constructed as cheaply as possible.

Frequently school authorities not wishing the trouble of constructing desks, desire to secure them ready made from some reliable business firm. Desks built by manufacturers, who make a specialty of such work, present many points of advantage; they are generally well designed and well made. Naturally their cost will be much greater than the cost of the simple desks mentioned above. The following parties make a specialty of manufacturing laboratory furniture, and it is advisable to obtain the catalogues of several houses while studying this matter:

The Kewaunee Manufacturing Co., (Catalogue No. 4), Kewaunee, Wis.

Leonard, Peterson & Co., 1240 Fullerton Ave., Chicago, Ill.

L. E. Knott Apparatus Co., Boston, Mass.

The *best design of an individual desk* found on the market is that of the Altaffer Individual Desk, which is designed and manufactured by L. B. Altaffer, 1445 Wyandot Ave., Cleveland, Ohio.

Laboratory Burners and Fuel Supply.

The *alcohol lamp* is so feeble that nearly all teachers agree in reporting it as *absolutely unsatisfactory*.

The small gasoline torch sold by all dealers (e. g. Central Scientific Co., Cat. No. 4449), is reported as quite satisfactory by some teachers and as unsatisfactory by others. The difference in opinion is probably due to the fact that the valves of the pump do not last longer than one session and then give much trouble; again, the gasoline must be filtered through a chamois bag, since otherwise the small gasoline vent in the burner is choked up frequently. Usually repairs have to be made by the teacher, hence with heavy laboratory work, and more than a few students, the loss of time caused by these burners *becomes prohibitive*.

The valve-less gasoline burner (e. g. Central Scientific Co., Cat. No. 4455) is rather feeble and exceedingly troublesome, hence it is not to be recommended.

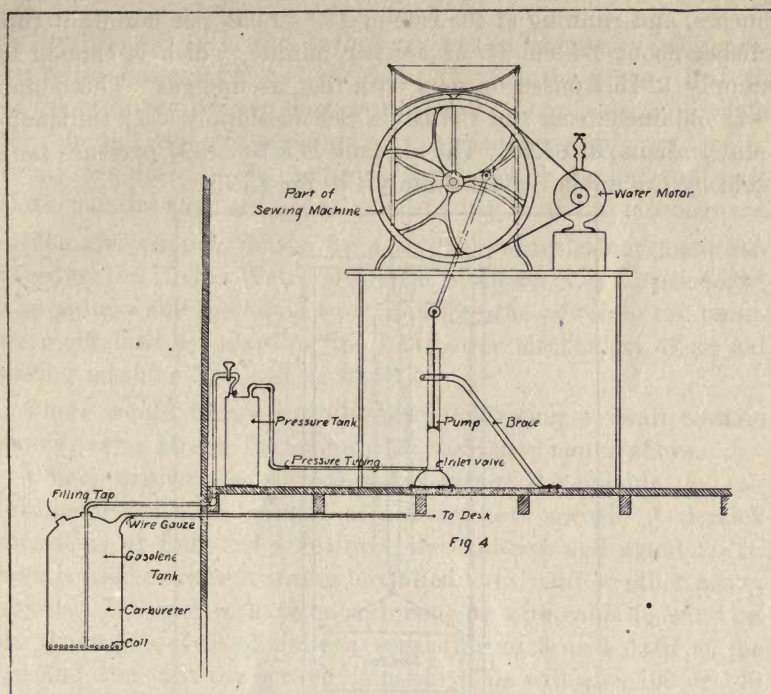
Acetylene gas may be burned in a Bunsen burner of special construction (Central Scientific Co., Cat. No. 4632). It is a very convenient and satisfactory source of heat, but its cost appears to be prohibitive.

Gasoline Gas Machines. There are at least three distinct machines of this sort on the market: the F-P Gas Machine, sold in this state by the Texas Lighting Co., Box 687, Dallas; the Standard Vacuum Gas Machine, manufactured by the Standard-Gillett Light Co., 9-11 West Michigan St., Chicago; and the Tirrill Gas Machine, manufactured by the Tirrill Gas Machine Ltg. Co., 509 Fifth Ave., New York. See also the Home-Made Machine farther on.

The smallest unit of the F-P Gas Machine is for five burners, and the prices up to the fifteen burner size are much smaller than the price of the smallest unit (twenty-five burners) of the cheaper one of the other machines; but for a larger number of burners the others are at least as cheap, when all is considered, and they are certainly to be preferred. The writer has not seen the F-P machine in operation, but he really doubts that it is suitable for chemical laboratories because the burners are permanently attached to the supply pipes, and the number of burners can be increased only by having the makers put in the additional equipment.

The two other gas machines, the Tirrill and the Standard Vacuum, are practically of the same type, and of quite a different type from the F-P machine. Of the Tirrell machines there are several now in successful operation within the state. Both machines are furnished in several sizes, the smallest of which furnishes gas for 25 lights. As quoted to the writer, the price of the Standard-Vacuum Machine is considerably less than the price of the Tirrill. Attention must be called to the fact that the prices for these machines are f. o. b. factories, and that the cost of their installation is considerable.

Although the purchase of one of these machines may appear to be an unduly large expenditure for this one need, namely, that of fuel, yet it must not be forgotten that these machines furnish the best fuel for laboratory purposes and save on other expenses, first of all on the cost of the burners, and second in saving the time which is otherwise lost by both the teacher and the pupil in keeping the other kinds of lamps in operation. Hence for a class of 20 or more a machine of this sort is much more economical than a number of gasoline blast lamps plus the cost of time for their operation. But even for smaller classes,



A Home-Made Gasoline Gas Machine.

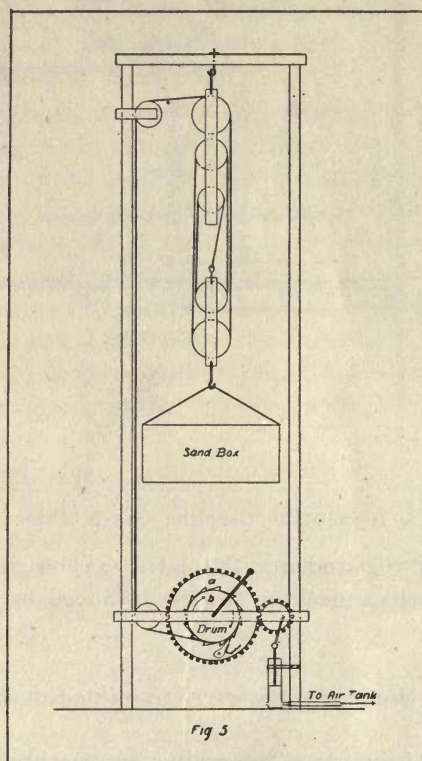
for instance, of 10 students, the extra expense involved in the purchase of such a machine will be balanced by the saving in several years.

A Home-Made, Inexpensive, Gasoline-Gas Machine.

A home-made, inexpensive gasoline-gas machine was installed in the Mexia High School last summer and it has been in successful operation during the fall. Mr. A. G. Koenig, who designed and installed it, has kindly furnished the following description of it:

The gasoline gas machine now in use in the Mexia High School consists of four parts: 1. The pump. 2. The air tank. 3. The carburetor and gasoline tank. 4. The source of power. These are assembled as shown in Fig. 4. The pump used is one having an internal diameter of 1.12 inches, the stroke is about 4

inches, and running at the rate of 150 strokes per minute it furnishes about 1-4 cu. ft. of air per minute, which is enough to supply 12-15 Bunsen burners with the gasoline gas. This pump was obtained from the Columbia School Supply Co., Indianapolis, Indiana, at \$4.90. The air tank is a No. 1827 pressure tank sold by Columbia School Supply Co. for \$3.75.



Mechanical Power for Gas Machine.

The carburetor consists of a five gallon oil can which can be bought for \$0.60, and about 3 ft. of 1-4 inch lead tube closed at one end and coiled up in the bottom of the can as shown in the diagram. A large number of very fine holes were made in this lead coil by means of a sewing needle or a fine awl. The air passing through these numerous holes in the coil at the bottom of the gasoline becomes saturated with gasoline vapor. It is advisable to place two rolls of wire gauze in the iron pipes con-

veying the gas from this carbureter to the laboratory; this prevents the flashing back of the flame through the pipes. Instead of the ordinary Bunsen burner it is advisable to use one in which the air and the gas supplies may be independently and accurately regulated. The expense for gasoline during the past three months has been about one cent per hour for ten burners.

The air pump is driven by a six-inch water motor manufactured by the Divine Water Motor Co. of Utica, N. Y., price \$5.00. The pulleys and mechanism for applying the power to the pump were obtained by adapting the foot-power mechanism of an old sewing machine at a cost of \$2.00.

There would be no difficulty in substituting a small electric motor (say a strong fan motor) for the water motor above.

Where neither electric nor water power is available the following device may be used as a source of power: A derrick consisting of four 2x4 scantlings, well braced, and about 20 ft. high is used to raise a strong box filled with sand or other heavy objects. In order to have enough rope or wire cable to wind on the drum, a system of pulleys, consisting of 3 or 4 fixed at the top and 2 or 3 at the bottom, is used. This will give 100 or 140 ft. of rope to be wound on the drum. The mechanism of the whole device is shown in Fig. 5.

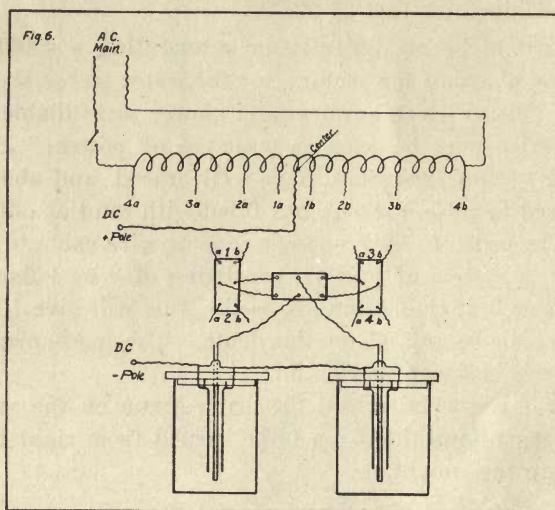
The large cog-wheel *a* and the drum *b* run on the same axle. The ratchet *c* allows the drum to be turned from right to left in winding up the weight.

Source of Current for Electrolytic Work.

As a rule, electrolytic work requires a *direct* current of low *voltage*. At present this is obtained either from battery cells (dry cells), or from high voltage (110 volts) direct-current lighting systems by wasting nine-tenths of the energy. Both of these sources are inefficient, uneconomical, or troublesome, and it is desirable to secure a better source. This is obtained by *transforming* and rectifying the alternating lighting current which is found even in the smallest towns. Since no really first-class, modern apparatus for this purpose appears to be in the market, the following detailed description for securing such is here given.

Two distinct rectifying sets are described below. The first

is the best design for this purpose that can be made at present. It naturally costs more than the second, but even its cost is much less than the cost of anything offered in the market that could take its place. In both sets the *electrolytic jars* or *rectifier proper* can be easily assembled at home, as well as, and perhaps better than, it is done by the manufacturers; and at much less



A Modern Current Rectifier.

cost. The second represents the design followed by many manufacturers; a combination of a lamp-bank rheostat with the rectifier proper. The particular form here described costs only a trifle, and is *very easily constructed*.

(1) A current rectifier for 10 amperes direct current and any pressure up to 45 volts. This apparatus consists of the following:

(a) A specially designed transformer (price, \$13.00, f. o. b., St. Louis, Mo., Moloney Electric Co.). For details see p. 116.

(b) Two electrolytic jars (or one double-jar); approximate cost of material \$2.50; for details of construction, see p. 114.

(c) Three double-pole, double-throw switches (smallest Bryant Baby Knife switches on porcelain bases); secure locally or from Western Electric Co., approximate price, \$1.50.

(d) Two binding posts, 30 cts.

(e) One rheostat with approximately 10 ohms resistance and a current capacity of 10 amperes on at least a part of the rheostat. A Ward-Leonard Co. (Bronxville, N. Y.) Field Rheostat, FF 810, price \$5.80, or FF 820, price \$7.30, will be found to be very serviceable, and probably cheaper than other equivalent rheostats. The whole is to be assembled as shown in Fig. 6.

This apparatus operates extremely economically and admits of convenient regulation to any desired current or voltage. Both alterations of the alternating current are "rectified," making the direct current obtained practically *continuous*. *It is just as serviceable, and cheaper than a storage battery.* The nearest approach to it in the market is a four-jar rectifier (see page 119 combined with a lamb-bank rheostat, the cost of which is as much, if not more, than this, while it is not nearly as serviceable.

(2) A cheap small current rectifier. This apparatus is designed for a maximum of about 3 amperes, at 4 to 8 volts; the total cost of material for the apparatus is about \$3.50, and it can be made, i. e. assembled, in one or two hours. It will be found *much cheaper, more convenient, and more serviceable than primary batteries or dry cells.* The details for construction are given on page 118. Only one alternation of the alternating current is rectified, giving a pulsating current. However it is just as serviceable for electrolytic work as a continuous current.

The details for construction, action, and operation of rectifying sets are given at length on page 113.

SUPPLIES.

These lists of apparatus and chemicals include all that is needed for the work given in most text-books, and also for the work outlined in this paper, with the possible omission

of a few things which are easily obtained locally. Naturally, special experiments given in particular laboratory manuals are not provided for here, but since nearly all modern laboratory manuals give lists of the apparatus and chemicals needed, the extra material may be easily added. This list calls for somewhat larger quantities than those given in the manuals, but these larger quantities will be found very desirable.

In ordering supplies, an individual set should be secured for the teacher. Of the perishable material, half as much again as is necessary for the individual sets should be secured.

Individual Outfit.

1 Burner, either a Bunsen burner with 2 ft. of rubber tubing of best white rubber wrapped in cloth, diam. 11x6.	\$.50
or a gasoline torch (\$2.75.)	
1 Wing top (for Bunsen burner only).....	.10
1 Iron stand (24 inches) with three rings.....	.65
1 Iron burette-clamp and clamp-holder.....	.35
1 Burette 50 cc. with glass stop-cock (See Note 1).....	.70
1 Deflagrating spoon.....	.03
1 Piece of wire gauze, iron, heavily tinned, 5"x5".....	.07
1 Piece of asbestos board 5"x5" $\frac{1}{8}$ " thick (See Note 2)..	.02
2 Porcelain evaporating dishes: 1-3"; 1-5".....	.25
1 Small mortar and pestle of wedgewood, $3\frac{1}{2}$ to 4" in diameter	.40
1 Wash bottle (1000 cc.) with rubber stopper and glass tube fittings (See Note 3).....	.30
1 Nest of 4 beakers—50 cc., 150 cc., 250 cc., 400 cc.....	.40
1 Crucible Tongs	.50
1 Porcelain Crucible with lid, 18cc.....	.20
1 Clay-covered Iron Triangle.....	.03
2 or 3 Wide Mouth (or salt mouth) Bottles, 500 cc. (See Note 7).....	.20
2 or 3 Pieces of Window Glass, 4x4.....	.10

12 Test-tubes, 16-160 mm. .01 each.....	.12
1 Test-tube rack with pins.....	.28
1 Dropping Funnel with cylindrical, open top, 50 cc. capacity; stem length, 12 cm., diam. 6 mm.....	.40
1 Test-tube Holder (wire).....	.08
2 Glass Funnels: 2½'' and 3½'' respectively.....	.25
1 Test-tube Brush.....	.03
1 Wooden Funnel Stand for 2 funnels.....	.60
1 Triangular File (See Note 4).....	.05
1 Pneumatic Trough (See Note 5).....	.35
½ Pack (50) Filter papers 12 ½ cm. (5'').....	.10
1 Thistle Top Funnel tube, length 12 in., diam. of stem, 6mm.....	.03
1 Graduate, or one open top, graduated, cylinder, 100 cc..	.15
2 Erlenmeyer Flasks <i>with two hole rubber stoppers</i> to fit..	.40
1 flask-250 cc. with No. 3 or No. 4 rubber stopper: 1 flask-400 cc. with No. 4 or No. 5 rubber stopper (See note 6).	
2 Watch Glasses: 3'' and 5''.....	.20
3 ft. Glass Rodding, diameter 3mm., per lb.....	.50
10 ft. Glass Tubing, diameters 6x4 mm. or 5x3 mm., per lb.	.12
2 Hard Glass Test-Tubes of best hard glass, 18x160 mm...	.20
1 Porcelain spoon and spatula, length 10 cm., width of spoon, 14mm.....	.10
1 Piece of platinum wire, No. 28 B. & S. gage, length 5 cm. fused into the end of a piece of glass <i>rodding</i>	.16
1 Piece of black rubber tubing, length 12 inches, diameters 7x4 mm.....	.20

Notes on Student's Outfit.

(1) All ground glass stop-cocks should be taken out and tied on the apparatus when not in actual use.

(2) This piece of asbestos board is to be used in place of a water bath, and for elementary work is much more serviceable. Should a water bath be absolutely required, then a large size beaker will probably serve the purpose. *No* water bath is included in the list.

(3) Buy flat bottom, 1 liter, Bohemian flasks, cover the necks with cord, secure two hole rubber stoppers to *fit mouth of flask* (probably No. 6 rubber stopper), and during one of the earliest

laboratory periods teach the student how to make the glass-tube fittings.

(4) Purchase the triangular files by the dozen from a local hardware dealer.

(5) Buy small dish pans and paint these inside and out with asphalt paint; they will serve the purpose just as well as the more expensive regular pneumatic troughs.

(6) These flasks with the rubber stoppers serve as generating flasks, as gas wash bottles, etc.

(7) Almost any sort of a wide-mouth bottle (pickle bottle, fruit-jars, etc.) may serve for these bottles. The pieces of glass to be used as covers may be supplied by a local dealer.

(8) All rubber ware, i. e. stoppers and tubing, when not in actual use, should be disconnected, or tubes drawn out of stoppers, etc., and should be laid away in the dark, since light decomposes rubber rapidly.

General Laboratory Supplies for a Class of Twelve Students.

One blast lamp and foot bellows for gas (or a large gasoline torch.) There is no foot blower listed by American dealers that is as satisfactory and as lasting as the one listed by Dr. Bender and Dr. Hobein, Munich, Germany, Cat. No. 2201. The blower, No. 2 diameter 30 cm. is listed at 40 mark (\$10.00) and is quite satisfactory. No. 2245 in the same catalogue gives an extremely useful blast lamp (list price \$3.00) to be used with the blower. A blast lamp outfit is *not absolutely necessary* but is extremely desirable. By its means a teacher can make many pieces of simple glass apparatus, which otherwise is impossible. Simple glass-blowing can and should be mastered by every teacher of chemistry. Helpful hints on glass-blowing are found in the booklet: Shenstone, "The Methods of Glassblowing." (Longmans, Green & Co.)

Six sets of eight desk reagent bottles, with labels blown in the glass, capacity 500 cc.; with ground glass stoppers (except where marked). (One set for every two students working simultaneously and one extra set for the teacher).

Concentrated Sulphuric Acid.....	} \$10.92 for 6 sets
Dilute Sulphuric Acid.....	
Concentrated Hydrochloric Acid.....	
Dilute Hydrochloric Acid.....	
Concentrated Nitric Acid.....	
Dilute Nitric Acid.....	
Ammonium Hydroxide.....	
Sodium Hydroxide (rubber stopper).....	

50 to 75 solution bottles for the general side shelf in the laboratory; capacity, 500 cc., ground glass stoppers. To be labeled as desired, with paper labels.

For this purpose secure one or two label books (.70). The labels should be pasted carefully so that the paper will not "pucker up", the "gum" allowed to dry thoroughly, and then a layer of melted paraffin spread with a brush over the label, beyond the edges.

The salt mouth bottles for the general supply shelf probably need not be purchased separately, as the bottles in which salts are shipped serve well for this purpose.

Balances for ordinary weighing, capacity 1000 grams. A very desirable balance is the Laboratory Balance, Cat. No. L 6010 (list price, \$12.00), manufactured by Wm. Gaertner & Co., 5345 Lake Ave., Chicago. One such balance will suffice for 10 students.

1 set of weights, third class, 10 grams up to 2 Kg. (This balance has a sliding weight for the smaller weight down to 0.1 gram).

Water Still. If running water is at hand, a small automatic still (for instance, a *Jewell*) should be installed (one of capacity $\frac{1}{2}$ gallon per hour will be ample for 10 to 20 students). If the laboratory is not supplied with gas, a small (1 hole) blue-flame kerosene stove should be bought, and placed in position to heat the still. If running water is *not* available in the laboratory, a (5 gallon) copper retort (tin lined) and a block tin worm condenser may be used. It can be heated with a Fletcher radial gas burner or with a kerosene blue-flame stove. Either distilling apparatus will cost slightly above \$20.00. Two 5 gallon stout

glass bottles for collecting and storing distilled water should be secured (\$3.00), and one of them should be fitted with a glass syphon and pinch cock for drawing the water.

<i>Two Acid Pots</i> , stone ware, 3 gallons capacity, for diluting acids		\$2.00
3 Sets of 3 smallest sizes of cork borers		1.50
1 Set of 6 smallest sizes of cork borers.....		.80
3 Kipps automatic gas generators, capacity 1 pt., with a broad base having a "tubular" in it for emptying the spent acid. For convenience in "filling," the tubular in the middle "globe" should be of 3½ cm. inner diameter		9.00
100 ft. Tubing of soft glass, for bending, etc., diam. 9½x12½ mm.....		.85
30 ft. Tubing of soft glass, for bending, etc., diam. 14x18		.50
100 ft. Tubing of soft glass, for bending, etc., 5mm. outer diam.		.42
100 ft. Tubing of soft glass, for bending, etc., diam. 6x4...		.50
100 ft. Tubing of soft glass, for bending, etc., diam. 7x5...		.45
30 ft. soft glass rodding, 3 mm. diameter.....		.15
10 ft. soft glass rodding, 6 mm. diameter.....		.10
6 Thermometers, solid stem, graduated up to 360° C.....		7.50
12 Calcium Chloride U-tubes, 6 in. with side tubes attached		.72
12 Mohr's pinchcock clamps.....		.98
12 Screw clamps for compressing rubber tubing.....		3.00
6 Lead dishes, 3 inches in diameter.....		.75
12 Distilling flasks, 350 to 400 cc. with effluent tube half-way up the neck. They are to be used in place of retorts. For a condenser, use a piece of soft glass tubing diameters 10x13 mm. Fit this over the side tube of the distilling flask by wrapping a strip of asbestos around the latter. When necessary place a wet towel around the condensing tube, or let this extend into a receiving flask submerged in water. To apply heat, place the distilling flask over a hole (1½ to 2 inches diameter) in a piece of asbestos board which is 5x5 inches in size. This prevents the flame from striking portions of the flask above the level of the liquid in it.....		1.50

2 Large porcelain spoons.....	.50
1 Measuring cylinder, 1000 cc. open top.....	.50
4 Volumetric flasks without glass stoppers:	
1—1000 cc.....	.35
1— 500 cc.....	.25
1— 250 cc.....	.20
1— 100 cc.....	.16
1 Measuring cylinder, 1000 cc. glass stoppered.....	.63
2 10 cc. Pipettes.....	.10
2 25 cc. Pipettes.....	.14
2 Glass funnels, 6 inches diameter.....	.40
4 Porcelain evaporating dishes, ordinary form:	
1—5 inch diam.....	.20
1—6 inch diam.....	.23
1—7 inch diam.....	.30
1—8 inch diam.....	.38
2 Large porous cups and 1 rubber stopper to fit these—to be used for the demonstration of gaseous diffusion. A bell jar can probably be borrowed from the physical laboratory for this purpose.	
Beakers of best Bohemian glass:	
3 nests each, about 600 cc., 800 cc. and 1000 cc.....	1.80
1 Package of 100 circular filter papers, 12 inch diam.....	.16
1 Glass tube cutter (inside cutter) with extra cutting wheels. (E. & A. Cat. 3684).....	1.25
1 filter Pump (Aspirator), of brass, medium size.....	1.50
“Picein” rubber cement, very convenient and useful for tightening joints and stoppers in apparatus. It is applied like sealing wax, but is preferable to the latter. Dealers can secure same from Fritz Koehler, Leipzig, Germany, 4 pieces at 15 cents apiece, approximately.....	.60
Cork stoppers, best quality—regular size.	

Diam. Small End.

100.....	.14.....	.38
“	.16.....	.50
“	.18.....	.60
“	.20.....	1.00
50.....	.22.....	.55
“	.24.....	.60

"	26	.65
"	28	.75
"	30	1.00
"	32	1.10

Rubber Tubing. Besides that necessary for the Bunsen burners and the black rubber tubing in the individual outfits, "for connections," secure the following:

Amount	Diam.	
20 ft.	13x 7 mm.	\$3.60
10 ft.	15x11 mm.	3.00

Special Apparatus for Quantitative Experiments.

Balance for quantitative work. This should be a balance in a glass case sensitive to a milligram. The capacity of the balance should be 300 grams, and a set of second (not third) class weights should be provided for it, and a supply of extra fractional weights should be on hand to replace any that may be lost. The balance shown in Eimer and Amend's catalogue No. 2122, and the set of weights No. 2198 (200 grams) are suitable for this purpose. In general, one such balance will be enough for seven or eight students only, hence if much of this work is to be done, two such balances should be provided for a class of twelve students. The price of such a balance and weights is about \$20.00 to \$25.00.

12 Porcelain combustion boats, 7x70 mm. or 10x60 mm.

Apparatus for the Section on Electrolysis, Ionisation, etc.

See special section on electrolysis and battery action. The following will be needed:

12 Porous cups, height 3 inches, diameter, 1½. They should be well protected against dust.

12 Rubber stoppers to fit these. Holes can be cut in them as needed. This is easily done by means of the ordinary cork borer if it is moistened with sodium hydroxide solution or with alcohol.

1 Drying Tower (E. & A. Cat. N. 3055), 18 inches high..\$.50
12 small carbon rods—i. e. arc light electrodes—can be secured locally.

- 4 Platinum electrodes each consisting of a thin piece of platinum foil 2x5 cm. to one end of which is welded a short piece of No. 22 platinum wire. The latter is fused into a four-inch piece of strong glass tubing (extending to the inside), having 7 mm. outer diameter, and for electrical connection, a little mercury is poured into the tube.\$5.00
- 2 Porcelain jars, with diameter 6" to 8," height 4" to 6" can be secured at a local dealer..... 40
- Ammeter, voltmeter, rheostat, etc., may probably be borrowed temporarily from the physical laboratory. If not, then at least one ammeter with two ranges (1 amp. and 10 amp.) and one voltmeter with two ranges (3 volts and 30 volts) should be bought. These instruments should be of fairly good grade, hence they should cost at least \$15.00 a piece.
- 1 Piece fairly stiff sheet copper (about the thickness used in lining bath tubs) 16x16 inches..... .50
- 1 Wooden "Conductivity" vessel made to order (see page 77 2.50
- 1/2 lb. Picric acid..... .75
- 1/2 lb Naphthalene c. p..... .25
- 1 Liter absolute alcohol..... .35

Apparatus for the Demonstration of Combining Volumes of Gases

- 1 Hoffman Electrolysis apparatus; for details see page 53.\$6.00
- 1 Eudiometer Tube of special design (see page 53) for the explosion of hydrogen and oxygen..... 5.00
- 1 Hoffman apparatus tube to demonstrate the volume relations for ammonia—this tube to be modified as stated on page 55..... 3.00

Chemicals for a Class of Twelve Students.

- 5 lbs. Acetic Acid glacial.....\$1.02
- 18 lbs. Hydrochloric Acid c. p. in 6 lb. sealed bottles..... 2.16
- 7 lbs. Nitric Acid, c. p. in 7 lb. sealed bottles..... 1.00
- 2 lbs. Fuming Nitric Acid..... 1.10

1 lb. Oxalic Acid, commercial.....	.12
27 lbs. Sulphuric Acid, c. p. in 9 lb. sealed bottles.....	2.85
1 lb. Tartaric Acid.....	.64
2 qt. 95% Alcohol.....	.50
1 lb. Ammonium Carbonate.....	.25
$\frac{1}{4}$ lb. Ammonium Oxalate.....	.12
1 lb. Potassic Alum.....	.08
$\frac{1}{4}$ lb. Aluminium, filings or shavings.....	.35
5 lbs. Ammonium Chloride, pure.....	.90
$\frac{1}{4}$ lb. Ammonium Sulphate.....	.10
16 lbs. Ammonium Hydroxide in 4 lb bottles.....	2.80
3 lbs. Ammonium Nitrate.....	.70
8 ozs. Antimony, powdered.....	.25
4 ozs. Arsenic Trioxide.....	.13
$\frac{1}{2}$ lb. Asbestos wool.....	.20
1 sq. yd. Thin Asbestos paper.....	.10
2 lbs. Barium Chloride, crystals.....	.25
$\frac{1}{4}$ lb. Barium Nitrate.....	.10
$\frac{1}{4}$ lb. Barium Peroxide.....	.35
1 oz. Bismuth.....	.30
1 lb. Borax	.18
$\frac{1}{2}$ lb. Bromine in 2 oz. bottles.....	.50
1 lb. Calcium Chloride, crystals.....	.35
2 lb. Calcium Chloride, anhydrous, fused.....	.80
1 lb. Cadmium Chloride.....	1.35
1 lb. Calcium Fluoride.....	.10
Calcium oxide, buy locally.....	
$\frac{1}{4}$ lb. Cadmium Nitrate.....	.40
1 lb. Calcium Nitrate.....	.80
$\frac{1}{2}$ lb. Chrome Alum, c. p.....	.25
1 lb. Chloroform, Squibb's.....	.90
5 lbs. Calcium Sulphate, Plaster of Paris.....	.15
1 lb. Carbon Bisulphide.....	.16
10 lbs. Calcium Carbonate, marble chips.....	.50
2 lbs. Bone Black.....	.15
1 oz. Cobalt Nitrate.....	.15
1 doz. sticks Willow Charcoal.....	.36
1 lb. Thin Copper Foil.....	.50
2 lbs. Copper Turnings, fine.....	.40

2 lbs. Copper Nitrate.....	.65
1 lb. Copper Oxide, black powdered.....	.75
5 lbs. Copper Sulphate.....	.50
2 oz. Iodine, pure, crystals.....	.75
2 lbs. Ferrous Sulphate in 2 oz. bottles.....	.40
1 lb. Ferric Chloride.....	.12
2 lbs. Iron Filings.....	.10
5 lbs. Iron Sulphide.....	.35
2 lbs. Sheet Lead.....	.20
1 lb. Lead Nitrate.....	.18
1 lb. Litharge	.12
1 lb. Lead Peroxide.....	.36
1 oz. Litmus	.10
1 Quire Litmus paper, red.....	.50
1 Quire Litmus paper, blue.....	.50
2 ozs. Magnesium ribbon.....	1.20
1 lb. Magnesium Chloride.....	.12
1 lb. Manganese Chloride.....	.50
2 lbs. Manganese Dioxide, powdered.....	.85
3 lbs. Manganese Dioxide, in lumps.....	1.00
5 lbs. Mercury.....	4.00
$\frac{1}{4}$ lb. Mercuric Chloride.....	.40
8 ozs. Red Oxide of Mercury.....	.80
$\frac{1}{4}$ lb. Mercurous Nitrate.....	.40
1 lb. Nickel Chloride.....	.70
1 oz. Phenolphthalein	.10
5 lbs. Paraffin	.75
1 oz. Red Phosphorus.....	.15
$\frac{1}{4}$ lb. Yellow Phosphorus.....	.25
$\frac{1}{4}$ oz. Potassium.....	.45
1 lb. Potassium Bromide.....	.20
2 lbs. Potassium Hydroxide, sticks.....	.60
1 lb. Potassium Carbonate.....	.16
5 lbs. Potassium Chlorate, crystals.....	.70
1 lb. Antimony and Potassium Tartrate.....	.60
2 lbs. Potassium Bichromate.....	.32
$\frac{1}{2}$ lb. Potassium Ferrocyanide.....	.30

1/2 lb. Potassium Ferrieyanide.....	.30
1/2 lb. Potassium Iodide	2.40
1 lb. Potassium Nitrate.....	.15
1/4 lb. Potassium Sulphocyanate.....	.15
3 lbs. Potassium Chloride	.25
1/2 lb. Potassium Permanganate.....	.10
1 oz. Platinized Asbestos.....	5.00
1/2 oz. Silver foil.....	.60
1/2 lb. Silver nitrate.....	4.00
8 oz. Sodium.....	.70
2 lbs. Sodium carbonate crystals.....	.25
3 lbs. Caustic soda, sticks.....	.75
2 lbs. Sodium chloride, c. p.....	.50
1/4 lb. Sodium chlorate.....	.10
3 lbs. Sodium nitrate.....	.42
2 lbs. Sodium sulphate, crystals.....	.25
1 lb. Sodium, sulphide, crystals.....	.10
2 lbs. Sodium hyposulphite.....	.15
1 lb. Sodium phosphate.....	.10
2 lbs. Sodium bisulphite.....	.60
1/2 lb. Sodium acetate.....	.25
1 lb. Sodium peroxide, commercial.....	.70
1/2 lb. Stannous chloride.....	.25
1 oz. Strontium chloride.....	.10
5 lbs. Roll sulphur.....	.25
1 lb. Flowers of Sulphur.....	.08
3 lbs. Granulated zinc.....	.45
3 lb. Zinc in sticks, at least 6 in. in length.....	.25
5 lbs. Zinc in lumps or short, thick rods, 1 in. long, "for Kipps"	1.25
1 lb. Zinc sulphate.....	.25
1/2 lb. Zinc nitrate.....	.50
1 lb. Granulated tin.....	.55
2 ozs. Glass wool, fine.....	1.00

Cost of Supplies.

The prices given in the list of "Individual Supplies" are those that would be charged at Austin, hence they include freight.

The prices for the other supplies are to serve as *indications* merely; they have not been obtained by special quotation from a firm; and they do *not* include freight. For the latter about 10 to 15% must be added to the price.

The total cost of the supplies in this list, figured for a teacher and twelve students, and including freight is:

Individual Sets.....	\$182.96
General Supplies.....	53.75
Freight and Drayage, 12½% on the latter.....	6.50
Total	\$243.21

First Cost of Laboratory.

If all furniture is constructed as cheaply as possible, its cost will be approximately as follows:

12 lockers, for students, including plumbing.....	\$ 75.00
Reagent side shelves, shelves in store room, work table for teacher, and large sink.....	50.00
One draft hood.....	25.00
Chemical supplies.....	250.00
Total	\$400.00

Cost of Maintenance.

The least annual cost of maintenance is about \$8.00 to \$12.00 per student. Of this amount from \$2.50 to \$5.00 may be borne by the student, and the remainder by the Board. In many Texas High Schools, the Board's appropriation is as much as \$10.00 per student annually.

PART II: PLAN AND CONDUCT OF THE COURSE.

Main Object of Course.

The question to be considered here is: for which set of students shall the course be designed—those who end their school days with the high school, or those who go to college? The writer believes that the course should be *designed primarily* for the student who will *not continue* his studies beyond the high school; and the following plan and suggestions have been made with that in view. Furthermore, the writer believes that such a course will also be the best introduction to the subject for the student who continues the study of chemistry in college.

For this purpose the course should consist mainly of a presentation of experimental facts connected by the least amount of theory necessary to aid in their study. Emphasis should constantly be given to the recognition and retention of the experimental facts themselves. Furthermore an essential part of the course should deal with facts with which the student will be concerned later in life.

Which first, Chemistry or Physics?

The question whether chemistry or physics shall precede in the high school course is a matter that should be left entirely to the high school people. Read in this connection, *The Teaching of Chemistry and Physics*, pages 29 to 37. Attention should be called to the fact that the progress of the class will naturally be slower if chemistry is given earlier in the course than if it is given during the last year. Furthermore, one topic must be given special consideration, that is the topic of battery cells and electrolysis. This subject should be given in the chemistry course rather than in the physics course because it is essentially a chemical subject. It cannot be presented without involving chemical considerations, and elementary chemistry cannot be presented properly without it (see page 81 et. seq.).

Time Allowance for the Course.

The time given to chemistry in the high school should be three recitation periods of 45 minutes and two laboratory periods of at

least double that time a week. Each double period for laboratory work should be uninterrupted, that is, there should be two consecutive periods. There should be no difficulty in arranging the schedule for this, because a study period may be placed just before or after the recitation period, and on days on which laboratory work is taken, the recitation period together with the study period then form the laboratory period. This arrangement is desirable for another reason: it frequently happens that the ratio between laboratory and recitation work temporarily should be changed in either way possible, because at times the recitation work requires more attention and at others the laboratory work. The arrangement of the schedule here suggested permits of such a change, and hence is highly desirable.

Time Allowance for Preparation of Laboratory and Lecture Table Experiments.

Considerable time must be spent by the teacher in the preparation of special apparatus, in the preparations of solutions for the laboratory, and in the preparation of lecture table experiments. In addition to this there is the unavoidable labor of putting away the apparatus properly after use, and arranging the side shelf reagents, etc., after periods of laboratory work. As a rule superintendents and principals either do not realize how time consuming this work is or they are disposed to minimize its importance. Some will consider that this is essentially janitorial work, for which some cheap help might either be hired, or which if done by the teacher does not deserve any consideration as a part of his teaching work. This is utterly wrong. The teacher of chemistry should be allowed time enough to do such work properly, and such time should be a part of his teaching time, and not of hours after school work. If the suggestion made above for the scheduling of the laboratory and of the recitation work is followed, then in addition to the double period each day, kept open throughout the week for chemistry, there should be set aside another period each day for the teacher to look after the equipment. In this way he will have two periods three times a week and one period twice a week, a total of eight periods to prepare his experimental work. This is not excessive and will be very profitable to the school. If classes are large, the teacher must be

given, in addition, assistants to aid him in cleaning up the laboratory, etc. Such service can be secured very reasonably from some member of the chemistry class. The cleaning up should not be done without the direction of the teacher, because he should know where everything is placed.

However, such time allowance, etc., puts a responsibility upon the teacher which he must not shirk. A chemistry teacher must be "up and about" doing experimental work with his own hands. All "sitting back" and letting the students only do experimental work or ordering around some assistant to prepare experiments and reagents is indictative of a person unfit to be a teacher of science. He must be the *chief experimenter*; he must be possessed of a desire to do experimental work, and with his own hands attend to matters in the laboratory.

Manner of Conducting Laboratory Work.

Beginners in chemistry are frequently inclined to do their laboratory work in a mechanical manner merely, and the newness of the subject inclines them also to take too much time in the performance of experiments. Hence the teacher must use a method of conducting laboratory work which will keep the mental aspect of the subject constantly before them and force them to do their work expeditiously. This may be done by conducting the laboratory work in a manner similar to that of conducting class work. Thus at suitable times during the periods of laboratory work, the teacher should discuss with the whole class the chemical changes dealt with, and drill his pupils on all related matter. He should strive by all means possible to enliven the activity of laboratory practice, realizing that it is in the laboratory rather than in the lecture room that chemistry may be taught and learned.

In the treatment of the experimental material the aim should be to teach the common properties of substances in the *broadest manner*. Thus in connection with the preparation of oxygen, the student should not be taught that "*the laboratory method for the preparation of oxygen consists of heating a mixture of potassium chlorate and manganese dioxide,*" but he should be taught that "there are some compounds of oxygen that decompose when exposed to a sufficiently high temperature and yield

oxygen as one of the products of decomposition. Of course not all substances break up in the temperatures at our command, yet quite a number do so, among which may be mentioned lead peroxide, mercuric oxide, potassium chlorate, potassium nitrate, etc. These should be remembered as substances not only rich in oxygen but that yield it up when exposed to a moderately high temperature obtainable in the laboratory." By this means the student has not only learned how oxygen may be prepared, but he has learned to some extent the behavior of oxygen compounds exposed to different temperatures. The information thus conveyed is farther reaching and more general than the mere *preparation of oxygen*. This effort to teach something of *general* bearing in connection with the specific information conveyed should be made wherever it is possible. In the Outline for a Course which is given farther on in this paper, this principle is applied rather carefully, as may be noticed by looking at the first few topics given.

In the teaching of experimental work it should be emphasized that an experiment is valuable *only if it is clearly remembered* by the student. There exists a mistaken notion that experimental work is intended primarily to train the student in the performance of such work somewhat for the same reason that training is necessary in carpenter work, metal work, etc. Such training is only an *incidental* object which is attained without particular effort if the work is done properly. The *main* object is to learn something about the substances, and the whole value of the work is lost if the experiment is forgotten. It is practically true that the value of the course is proportional to the accuracy with which everything, including details of a set of well selected experiments, *is remembered*.

Note on Quantitative Experiments.

Although quantitative experiments are usually introduced early in the course, yet it appears to the writer that it is inadvisable to do so. It is only when the student has a fairly extensive knowledge of chemical *phenomena* that he can appreciate *quantitative* relations. However, later in the course quantitative experiments are extremely valuable. In this connection see page 52.

The Note Book.

The first thing to insist upon in the note books is correct spelling and correct composition.

Next, all attempts to use blanks, to be filled in, or detailed "patterns," to be followed for the writing up of experiments, are to be discouraged. A few general directions are naturally necessary, but any approach to a "formula" to be used in writing up experiments should be decidedly discouraged.

The first general direction given to students for writing up their notes is to call attention to the fact that the name of the book should be indicative of its use; it should be a *note book*, and *not* a re-written text book or a book of essays. If in performing the experiments, printed directions have been followed, then these *directions should not be copied into the note book*. Of course such directions should be definitely referred to in some such manner as this: "Experiment (number) ——— page ——— in Smith's *Laboratory Manual* was performed in full (or "with the following exceptions"). If *additional* experiments in which printed directions were not followed have been performed by the student, then these may be written out in full.

In the note book the experiments may be numbered merely for the sake of indicating the order. However, the most important thing is to *devise for each experiment a heading* which states clearly the nature and object of the experiment. Considerable time may be spent in thought in order to devise headings that may be properly significant. A suitable heading is just as important in itself as all the notes that follow, and the devising of proper headings is one of the best means to force the student to realize what he is doing and what he is doing it for. The heading should not be merely a general term, and not more than one definite subject should be put under it. Thus it is absurd to give the heading "oxygen" to all the experiments given in connection with the study of that subject. Some of the experiments in connection with oxygen have been made with the view of demonstrating with what chemicals and with what operations it may be prepared. Other experiments may have been made with the view of preparing certain bottles

ful of the gas and showing some of its common properties. Hence it would be logical to head one experiment as follows: "To demonstrate with what substances and what operation oxygen may be prepared." The other experiment should be headed somewhat as follows: "The preparation and collection of oxygen gas and the demonstration of some of its properties."

The assignment of separate headings to distinct parts of experimental work should not be carried to the extreme. It is unnecessary to divide the last experiment discussed above into the following: *first*, the preparation and collection of the gas; *second*, the combustion of sulphur in oxygen, etc. The collection of several operations under one heading naturally requires thought, and during the first part of the course the student will have much difficulty with it, but the benefit gained both as far as grasping the subject is concerned and as far as training in composition is concerned will amply repay the time spent upon this work.

In the body of the "write up" the first thing to be given is a direct reference to the directions followed, e. g.: *Smith, Exp. ———, page ———, was performed in full*" (or with the following exceptions"), etc.

Then should follow a brief definite statement of what was actually noted or observed, which has not been given in the printed directions referred to above. For example, under the last heading mentioned would appear: sulphur, carbon, phosphorus and sodium after they were heated to their kindling temperatures burned vigorously in the oxygen gas. The product of combustion of carbon (carbon dioxide) is an invisible, inodorous gas; that of sulphur is an invisible gas with suffocating odor; that of phosphorus is a white solid and that of sodium is a white solid. Treated with water, these substances give solutions the first three of which turned blue litmus red and the last turned red litmus blue. The equations for the chemical actions which take place during combustion and afterwards on treatment of water are as follows: "."

The notes should always show the equations of the reactions in the experiment. Care must be taken by the teacher to see that they are written correctly.

How to Begin the Course.

Avoid *beginning* by giving definitions, for instance of *elements, compounds, chemistry, atoms, molecules*, etc. Starting with definitions reverses the attitude of the student towards the subject matter, and is really contrary to the natural method of learning a new subject. We teach a child a certain thing by mentioning its name and showing the thing, and perhaps speaking about it. This should be the mode of procedure in the introduction of this subject, which presents entirely new things to the student. Without attempting an explanation, the teacher should bring before the student the preparation and properties of several elements, oxygen, hydrogen, chlorine, etc., and tell him that *these* are elements. The formation of compounds, the various indications that *chemical* changes take place, are all shown incidentally in connection with these examples, and after a little while the student will be in position to form a definite notion as to what the terms, elements, compounds, chemical actions, etc., mean. The attempt to introduce the fundamental terms abstractly at the very beginning of the subject consumes a great deal of time and is rather unprofitable.

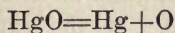
Introductory chapters concerning experimental examples of chemical change, the influence of heat, what is meant by a solution, exercises in weighing, measuring, etc., are also of questionable value. At best they are rather unfruitful and certainly uninteresting to the student. At the University of Texas they are omitted altogether, and the first exercise given is the study of the preparation of oxygen.

The Introduction of Symbols and of the Notions of Atoms and Molecules.

It is a mistake to speak constantly about atoms and molecules and treat them as something absolutely essential to chemistry. Students are *naturally* inclined to consider that atoms and molecules are the main things with which chemistry deals, hence a particular effort must be made against this inclination. Emphasize, at the beginning, that chemistry deals with *facts*, with things in this world as they are, and that the

notion of atoms and molecules is introduced merely for the purpose of forming a simple mental picture.

It is highly desirable to use chemical symbols and formulae as early and as extensively as possible. However, these are not necessarily dependent upon any *notion* of the existence of atoms and molecules. They are mere *short-hand expressions* for the *relations of the weights of substances which undergo chemical changes* (which relations are merely the results of quantitative experiments). When it has been determined that 108 parts by weight of mercuric oxide give 100 parts by weight of mercury, and 8 parts by weight of oxygen, then instead of stating this fact in the way that it was just stated, chemists state it by writing



and the student need only be told that this is a special short-hand system devised to represent or express the proportions of these substances, the symbols standing for certain numbers selected partly with an eye to simplicity or convenience. The symbols *need* not mean anything else, *and this is their fundamental meaning and main use*. Used in this simple sense, and for this definite purpose, symbols and equations should be introduced from the very beginning, and the fundamental (quantitative) significance should be constantly referred to by the teacher.

After some seven or eight chapters dealing with different kinds of chemical substances have been studied, then the teacher may point out briefly (1) that we imagine any one element made up of distinct little pieces all of the same weight (atoms), which weights show the same relation as the numbers represented by their symbols, and (2) that compounds are merely collections of clusters of the atoms of elements in such bunches as are indicated by the formulae of compounds. With this the student has been told all that he *need* be told for the present about the atomic and molecular structure of matter. All attempts to show how the atomic weight table was derived, all attempts to show that the atomic hypothesis rests upon the laws of *definite* and of *multiple* proportion will necessarily be philosophical, and it is questionable whether such attempts can be made profitably at this time. (In this connection see page 52.)

Nearly all text books consider fairly early the relations by volume of combining gases, which naturally leads to the presentation of Avogadro's Hypothesis and the determination of molecular weights. *These topics as constituents of the first two-thirds of the high school chemistry course are wholly undesirable.* They should not be given till later. If given early they tend to incline the student too much toward the theoretical aspects of the subject, an inclination which is altogether too great in most young students.

However, toward the end of the course such topics may be introduced, provided some of the experimental facts which show the relations by volume in which gases combine are actually presented. The emphasis is to be placed on the *presentation* of the *experiments*. If the experiments are not performed, then the whole topic should be omitted. In order that the teacher may not lose time in selecting simple, easily performed experiments for this purpose, a list of suitable experiments and apparatus is given in this paper (see page 53). However, in any case, the consideration given to these topics should not be too great.

The Introduction of Valence.

Say nothing about valence until several substances have been studied. In studying the first few substances give the students the formulae and equations and ask them to memorize these blindly. State merely that they are "short-hand" expressions for the relations of the weights of the substances involved. Note where and how valence is introduced in the "Outline of the Course," page 66.

The Fundamental Facts of Electrolysis and Its Introduction in the Course.

Do not give the electrolysis of dilute sulphuric acid, nor the electrolysis of any other substance as *examples of the decomposition of these substances*, and do not attempt to show by means of the relation of volumes of the gases obtained by electrolysis that water is composed of two volumes of hydrogen to

one volume of oxygen. The first fundamental and most important fact concerning electrolysis and the action of batteries is that *the chemical actions at the two poles are chemically in no wise related*. The actions at the two poles are also very complex and it is only under special conditions that the relation of the volumes of the gases produced by electrolysis is just the same as that in which they combine to form water. The latter fact can be demonstrated by showing that the two gases, mixed in the proper proportion and exploded, combine completely to form water, but it does not follow as a conclusion from the results obtained by electrolysis.

Electrolysis is a complex subject, and should not be treated in the early part of the course, but should be left to be presented in connection with other electrolytic phenomena so that the whole subject may be clearly and connectedly shown by experiment. An example of a modern treatment of this subject is presented at length in Part III, page 81.

General Discussion of an Outline for the First Part of an Introductory Course.

In the foregoing there has been presented (page 38) the *first* fundamental principle for the selection of topics in this course, which principle may be restated thus: make the course primarily a presentation of experimental facts and introduce theory to as slight an extent as possible.

The second fundamental principle governs the arrangement of the material. This should be arranged with the view of presenting *at first* only simple types of reactions, particularly reactions of hydration and metathesis, and deferring the presentation of the more complex types until the simpler types have received thorough consideration.

In the arrangement of subjects found in most text books neither the general types of reactions nor any classification according to the most prominent chemical properties of the substances have been considered. Thus, nitrogen and its compounds are invariably presented very early in the course, although the reactions between nitric acid and metals are always complex. They involve not only the simple reactions of hydra-

tion and double decomposition, but they also involve oxidation and reduction, and hence they should really be placed at a point farther on in the course. Again, many *simple* reactions which are presented by the interaction of salts of metals with common reagents, such as silver nitrate with hydrochloric acid, ferric chloride with sodium hydroxide, etc., are usually given at the end of the course, and yet they are the simplest kinds of chemical changes. As a result of the constant mixing up of all kinds of changes from the very beginning of the course, the student generally fails to grasp the fundamental principles in any, and instead of learning the properties of elements and compounds in their proper relation, that is, each as an example of a *general* fact, he tries to learn them all as isolated *specific* properties.

During the last thirty years our knowledge of reactions in aqueous solutions has increased so much that we may now be certain of the fundamental facts concerning them: it is definitely known that metathetical reactions depend upon the degree of ionic dissociation, while reactions involving oxidation and reduction depend upon additional facts, namely, those learned in the study of the galvanic batteries. The study of chemistry is immensely simplified by the application of this knowledge and a course arranged to set forth clearly these two fundamental types of chemical changes will enable the student to get a grasp of the chemically important properties of substances which is otherwise not so easily obtained. For this reason it is advised here to arrange the material according to our second fundamental principle; more particularly, it is advised to present at first only substances which undergo reactions of hydration and metathesis, and then to present later on reactions involving changes of valence.

Although the arrangement of the subject matter according to this principle is quite different from that found in textbooks, it will not be found difficult to follow this principle while using any good text. In order to make it convenient for teachers to do this, a detailed outline of the course prepared for this purpose is presented on pages 62 to 111. This outline is also issued separately so that it may be put into the hands of every student. The special experiments and directions which are

not found in text books have been written out fully. It is intended that this outline may be used together with any one of the text-books and laboratory manuals listed on page 60. The separate prints of the outline may be obtained from the author.

An examination of this *Outline of the Course* (see page 62) will show that in its first part it differs only slightly from the presentation usually given; but farther on in the presentation of the fundamental facts involved in metathetical reactions it differs radically from what is ordinarily found in texts. For details, see pages 73 to 80.

Since the reactions of the compounds of the metals involve nothing but metathetical reactions, these reactions are taken up next in order after the general discussion of the metathetical reaction.

How much of the reactions of the compounds of the metals should be taught in such a course as this? This has been, and possibly is still a great question. On one point, however, there is unanimity of opinion, and that is: *systematic qualitative analysis must not be attempted in a high school course.*

But this leaves us with the question: which of the reactions of the metals should be taught and how should they be taught? An examination of the laboratory manuals now on the market reveals great differences of opinion. Some manuals, such as Alexander Smith's and B. W. Peet's present the reactions found in the ordinary outline for qualitative analysis, the only difference between this arrangement and that used in qualitative analysis being that in this case the reactions are presented disconnectedly, and outlines for the systematic analytical procedure are deliberately omitted.

Other manuals such as the *Laboratory Exercises* by Brownlee and Others make no particular attempt to teach the reactions of the metals; but the latter book in Exercises 47 to 51 indicates that it is desirable to present certain features of this work. These features, which all authors are trying to present by their various methods, may be summed up as follows:

1. *To change* or transpose a metal from one of its compounds to another by means of the common reagents.
2. *To separate* the compound of a metal out of its mixture with compounds of other metals.

3. *To recognize or identify a certain metal in its compounds.*

The first object is illustrative of the methods of manufacture of many commercially important salts and compounds from the commercial sources of the metal. The second is particularly illustrative of the chemical purifying of substances. The object of the third is self-evident. There can be no question as to the desirability of including some of this work in the course and the only thing we need to consider is its extent and the method of treatment. With reference to its extent, that presented by *Brownlee and Others* is inadequate, while that presented by *Smith* and by *Peet* requires too much time. *Newell* uses the same method as *Smith* and as *Peet*, but he attempts to save time by lessening the number of exercises. This makes the work weak, if not utterly valueless. As a way out of these various difficulties, the following experimental procedure is suggested: take solutions of water soluble salts of the common metals and study their reactions with several general reagents. Then add special properties of some of the compounds of the metals such as the solubility of lead chloride in hot water, the hydrolysis of bismuth salts, the flame colorations, etc. Finally, add a list of exercises designed to make the student use these facts in the solution of special problems. The list of exercises given in the Outline is drawn up with this in view. These reactions of the metals give the student enough experience with the metathetical reaction to become well grounded in its fundamental facts.

Reactions involving oxidation and reduction—i. e., involving *changes* in valence—may be taken up next. These should be considered separately from hydration and metathetical changes because they involve entirely different fundamental facts. They should not be considered until *after* changes of hydration and metathetical reactions have been fully mastered, because they involve the latter.

The terms oxidation and reduction are rather misleading; many changes to be considered under this heading do not involve oxygen in any apparent manner. As the terms are now used, their direct significance is as follows: oxidation is an increase of positive electric charges on an element or radical (or the ab-

straction of negative electric charges); and reduction is the reverse electric change. The conception involved here can not be considered without a study of the action of battery cells, because the battery cell is, in a sense, *an experimental analysis* of oxidation and reduction reactions which presents the essential parts in their true relation. Any mixture in which oxidation and reduction take place represents merely the case in which the electric charges, instead of being transferred by means of conducting poles and their connecting wires, are transferred by direct contact between the two parts undergoing the changes in valence. Hence battery cells should be studied in this connection (see pages 87 to 91).

Following this presentation of the battery cell, oxidation and reduction may be introduced as given on pages 101 to 111.

The question now arises: How much of this last subject should be taught in a high school course? The *greatest danger* is to attempt *too much of it*. The work presented in the accompanying outline, extending through the oxidizing actions of nitric acid (to page 108) is *as much as should be given*. However after this, whenever in the presentation of descriptive matter connected with any topic there occur equations for oxidation and reduction actions, then the valences of the "valence-changing elements" or radicals should be plainly marked in the equations, as shown in the accompanying introduction to this subject (see page 106). By this means the equations will be made vastly more significant than without it.

Other Subjects That May Be Studied.

Beyond this point other subjects may be taken up in almost any order that may appear desirable.

The following subjects are worthy of consideration here:

- (a) elements not yet studied and their important compounds,
- (b) theoretical topics, such as the atomic and molecular theories,
- (c) topics of practical interest.

Under (a) the following subjects may be considered: first of all, the halogens (bromine, iodine, fluorine), if they have not

been previously studied; then any of the following in whatever order desired: phosphorus, arsenic, boron, manganese, chromium, and the oxy-halogen compounds. Only *a few of the most common experimental facts* should be considered. The classification of the elements according to the periodic system should be pointed out briefly.

Under (b) in connection with the atomic theory, the law of definite proportion and the law of *simple* multiple proportion may be considered. The law of combining weights (see Smith's College Chemistry, page 34) is probably too difficult and may just as well be omitted. At this period of the student's development, it would not be out of place to present the philosophical considerations, which, based upon these general facts, have built up the conception of the atomic structure of matter. However, *do not mix with this topic any consideration of the molecular theory based on Avogadro's Hypothesis*. This mistake is found in many texts, and must be guarded against. It is very desirable to give some quantitative experiments in connection with this topic, such as the determination of the hydrogen equivalents of metals, or the ratio in which metals combine with oxygen or with chlorine. In the choice of these experiments the teacher must exercise great care to select those only that are most easily *and accurately carried out*. The law of definite proportion may also be illustrated by the repeated titration of a base with an acid, using different quantities each time. Other suitable quantitative experiments are determinations of the solubilities of salts, the determination of water of crystallization in salts, etc. Good workable quantitative experiments are found in *Smith*, Chap. 7.

If the molecular theory, and Avogadro's Hypothesis are introduced, *several fundamental demonstrations should be made*. If such demonstrations cannot be made, then it is questionable whether the topic should be taken up at all. The following three experiments are probably the simplest and most workable that can be used in this connection, and those wishing to present this topic are strongly advised to perform these, at least.

Experiments Which Should Accompany the Study of Avogadro's Hypothesis.

(1). TO DETERMINE THE RATIO BY VOLUME IN WHICH HYDROGEN AND OXYGEN COMBINE TO FORM WATER.

For this purpose it is advisable to secure a specially made eudiometer tube constructed as follows: a very heavy glass tube of 50-60 cm. length and about 100 cc. capacity terminates in both ends in stout tubes with a small, almost capillary, bore, in which are fitted specially well ground stop-cocks. One of these terminal tubes should be bent twice at a right angle, its stop-cock should be placed about 1 inch from the open end of the capillary tube, and the tube should be calibrated, beginning at the stop-cock and marking every half cc. The tube should also be supplied with platinum wire electrodes placed in the wide tube near the end to which the bent capillary terminal is attached. It should be specified that the connecting loops of the platinum wire electrodes outside of the tube should be constructed so as not to be easily breakable. In addition to this tube there should be secured a Hoffmann electrolytic apparatus such as is shown in Eimer & Amend's catalog No. 3902. It should be specified, however, that the terminal tubes in which the stop-cocks are placed and through which the gases may be discharged, should have a small, almost capillary bore, and it should also be specified that the outer terminals of the electrodes should be specially well constructed so as to be practically unbreakable. To fill the eudiometer with hydrogen and oxygen it is first filled with water by suction, care being taken to fill the connecting tubes with liquid. The eudiometer is then placed beside the electrolytic apparatus, the bent end of the eudiometer tube is connected by means of a piece of rubber tubing with one of the gas-discharging tubes of the electrolytic apparatus, care being taken to bring the two ends of the glass tubes *in direct contact*. About 20 cc. of hydrogen and exactly half as much oxygen are thus transferred to the eudiometer. In measuring the gases in the eudiometer the differences in pressure due to the differences in the heights of the column of water in the eudiometer tubes may be neglected. The total vol-

ume of explosive mixture used should not exceed 25 cc. or at most 30 cc., since otherwise the explosion may shatter the tube. Before exploding the mixture the lower stop-cock is also closed. The lower end of the tube should extend into some water in a beaker, and after the explosion the lower stop-cock should be opened rather cautiously, so that the inrush of the water may not shatter the tube.

This apparatus is simple, and easy to handle, and by its means it is easy to demonstrate exactly that two volumes of hydrogen combine with one volume of oxygen. Even at the risk of being tedious the author cannot refrain from pointing out again that this fact cannot be demonstrated by means of the simple electrolysis of dilute sulphuric acid,—which is *erroneously* designated as the electrolysis of water. In this case, the production of the two gases is due to two entirely separate reactions, as has been pointed out elsewhere. Frequently the ratio of the volume of the hydrogen to the volume of the oxygen obtained by electrolysis is even greater than “two,” on account of the fact that not all of the reaction at the anode produces simple oxygen gas.

Since, for the same reason the electrolysis of hydrochloric acid cannot be considered to be a demonstration of the fact that one volume of hydrogen combines with one volume of chlorine to form hydrochloric acid, it appears to be desirable to demonstrate the direct union of hydrogen and chlorine. However, this reaction is so violent and so troublesome that it is not advisable to attempt it for classroom demonstration. Mere mention of the experimental result should be made in this connection.

(2) TO DEMONSTRATE THE RATIO BY VOLUME BETWEEN AMMONIA GAS AND THE NITROGEN OBTAINED FROM IT.

The next important point to demonstrate is to show the volume relation between original and resulting gases. For theoretical reasons it would be most advisable to use the example just referred to, by means of which it would be found that half a tube of hydrogen plus half a tube of chlorine would give a tube

full of hydrochloric acid gas. On account of the difficulty of performing this experiment it is inadvisable to use it as a demonstration. To demonstrate such a relation the following experiment should be performed. A tube full of dry ammonia gas should be treated with bromine (more specifically sodium hypobromite) by means of which all the hydrogen in the ammonia is abstracted and the nitrogen left free. This nitrogen, it will be found, occupies just one-half the volume occupied by the ammonia from which it was obtained. The Hoffman lecture apparatus tube for the demonstration of the volume relations in ammonia, which tube is shown in Eimer & Amend's catalog No 3904,—this tube should be obtained with an additional stop-cock attached to the closed end. The bore of the tube of this stop-cock should not be less than 2 or 3 mm. and the tube should be cleaned and dried by rinsing it first with water and afterwards with alcohol, and finally drawing air through it with the aid of the suction pump or by means of the foot pump, and at the same time warming the tube gently with the flame. It is clamped in a vertical position and filled with ammonia gas generated by boiling a concentrated solution of ammonia, and passing it through a drying tower filled with quicklime. The gas should be passed in at the top of the tube at a fairly good rate. After the tube has been filled with ammonia and closed, it can be laid aside until the time for the lecture. For the decomposition of the ammonia about 50 to 100 cc. of 20 per cent sodium hydroxide solution is treated with bromine until the solution has a straw yellow color. The mixture must be stirred vigorously while adding the bromine. This solution is added cautiously through the "funnel end," care being taken not to admit air. Finally the upper stop-cock is closed, and the lower stop-cock is opened while this end of the tube is dipped under water so that the latter may rise in the tube until the pressure inside the tube is equal to the atmospheric pressure.

This experiment is very easily performed, and gives the facts for the following argument. If the "tubeful" of gas contains " a " molecules, then the half "tubeful" contains " $a/2$ " molecules and since each one of the " a " molecules of ammonia contains at least one atoms of nitrogen, then there

must be at least "a" atoms of nitrogen in " $a/2$ " molecules of nitrogen, or at least two atoms in each molecule of nitrogen. The corresponding facts concerning hydrogen and chlorine producing hydrochloric acid gas and also hydrogen and oxygen producing *steam*, with the same argument, lead to the conclusion that the molecules of hydrogen, of chlorine, and of oxygen, each contain at least two atoms.

(3) TO DEMONSTRATE THE RATIO BY VOLUME BETWEEN OXYGEN AND THE SULPHUR DIOXIDE FORMED FROM IT.

For this purpose secure a piece of hard glass tubing about 24 inches long, closed at one end, and bent almost at right angle at a point about 5 inches from the closed end. Place a little sulphur in the closed end of the tube and fasten it there by melting the sulphur slightly. Then fill the tube partly with oxygen over mercury and place it with the short arm in a horizontal position. The action is started by gently heating the sulphur. Usually no gas will escape, and after the tube has been allowed to cool, the mercury will rise to the original position, which shows that the volume of the sulphur dioxide is equal to the volume of the original oxygen.

Chemical Information of Direct Economic Value in Texas.

Perhaps the most important additional topic to be added to the course is the presentation of subjects of practical interest. While the two preceding topics may be omitted entirely without any loss to the student, this last topic should not be omitted, because the major part of the direct value of the course to the high school student is obtained through the proper conveyance of accurate information on practical subjects with which he will have to deal more or less during the rest of his life. The value of such work and the desirability of presenting it in this course have never been doubted, but the results, obtained by many of the attempts made to present it, have frequently been worth very little, and occasionally they even serve to discredit both subject and teacher in the eyes of the general public. This back set, in general, is due to two causes: the first is the ignorance

of many teachers in the practical application of chemistry, and the second is a mistaken method of presenting it due to the zeal of the teacher to give this subject prominence. The subject is so important that proper methods of presentation and errors to be avoided should be given here.

Teachers just beginning to teach chemistry, whether college graduates or trained in other ways, are as a rule, not only ignorant of the practical applications of the subject, but their interest in the *pure science* inclines them to stress the pure, scientific subjects to the exclusion of the practical ones. If they continue to teach the subject, sooner or later they are impressed with the importance of the practical applications and in their effort to give this prominence, they may go to the other extreme. The writer has at hand some outlines prepared by prominent teachers of chemistry in various parts of the United States, in which an attempt is made to introduce practical subjects from the very beginning of the course. Long before the student has any knowledge of the simple chemical facts involved, these outlines present problems of water softening, the testing of coal gas for impurities, the solvent action of water on lead, the tests for organic matter in potable water, etc. Such information, though valuable in itself, becomes valueless when presented so early in the course that the student's lack of knowledge of pure chemical phenomena prevents a recognition of the simple *fundamental* phenomena involved. The mere mechanical testing of certain constituents of a potable water before the student has even learned or seen reactions between pure salt solutions, such as the reaction between sodium carbonate and calcium chloride, cannot be of much value either as an informational or as a training exercise.

Another mistake is the selection of applications which are nothing but empiric operations. Thus we find various attempts made to present the chemistry of cooking, the chemistry of cleaning, etc., under which headings many useful topics are offered. When, however, such subjects are presented by themselves, and particularly when the presentation crowds out fundamental chemical phenomena, they degrade the course into the mere "trying out" of a number of shop recipes.

A course filled with such material cannot give a well developed exposition of the subject. However, the entire neglect of these applications is probably just as great a mistake. The problem to be solved is to present enough of the fundamentals to give the student some sound training and then to add sufficient applications to give the subject "life." *The first thing necessary to accomplish this end is for the teacher to read and study in order to inform himself on practical subjects.* As a rule he will not get this information in college: he must secure it by informing himself in a general way,—as for instance, by reading books affording information on practical topics, by noting the operations and practices of practical men and artisans around him, etc. To be more definite, he should attempt to inform himself on such subjects as the following: soils, fertilizers, and irrigation waters; feed and food-stuffs; the technology of petroleum oils, tar oils, cotton seed oil, linseed oil, the manufacture of soap, paints, paper, etc. The following books will furnish him with a great deal of information on these subjects:

Soils and Fertilizers, by Harry Snyder (The Macmillan Co.).

Alkali Soils and Irrigation Waters, by G. S. Fraps, Bull.
130 of the Texas Agricultural Exp. Station.

Nature and Use of Commercial Fertilizers, G. S. Fraps, Bull.
112, Texas Agricultural Exp. Station.

Outlines of Industrial Chemistry, by F. H. Thorp (The MacMillan Co.).

Chemistry of Plant and Animal Life, Harry Snyder (The MacMillan Co.).

In the study of these subjects it will be found advisable to direct the attention to the following points:

The chemical analysis of a soil at present does not enable one to judge what ails an infertile soil; the value of fertilizers is determined by "pot" or "basket" experiments; the essential constituents of fertilizers are the nitrogen, potassium and phosphorus compounds in it; why rotation of crops is necessary; the Texas feed analysis law requires the determination of carbohydrates, fats, proteins, indigestible matter or crude fiber, water and mineral solids; what is meant by these divisions of food stuffs and how they are experimentally deter-

mined; the relative proportions of these constituents, in food or feedstuffs; the essential properties of the three classes of foodstuffs, that is, of carbohydrates, of fats, and of proteins; starch can be converted to dextrine and to glucose, and the latter can be fermented to alcohol; fats of vegetable or animal origin can be used to make soap; how cotton seed oil is made and refined; the essential difference between cotton seed oil and linseed oil, which makes the latter valuable to be used in paints; white lead is the most valuable pigment for paints and the best paints contain this together with just enough of some coloring matter to give the desired tints; the essential constituents of coal and their determination; how hard waters may be softened cheaply for boiler purposes; how potable waters may be purified in case they are unsanitary; how kerosene, gasoline, paraffin, etc., are obtained from crude petroleum and how they may be refined; what is meant by the flash point of kerosene; oils obtained by the distillation of tar are chemically more active than petroleum oils and hence form the starting point of the many synthetic dyes and drugs; oils obtained from tar, in contradistinction to the oils obtained from petroleum, are germicidal, and hence the cheapest portions of the tar distillates are used for wood preservation; how cement is made; the chemistry of photography and of blue-printing, etc.

If the teacher has no knowledge of organic chemistry, he should study certain parts of Remsen's *Organic Chemistry*. This book, though small, will give him the necessary scientific foundation for the parts in the foregoing list which deal with organic chemistry.

The introduction of this information into the course is rather difficult at present on account of the absence of a suitable textbook. The best text-books on the market today present a good introduction to the science, but do not present in any adequate manner such "practical" information as above indicated.

It is difficult to say how much of this can be done in a high school course, but it is certain that none will be done unless the teacher is actively engaged in increasing his own knowledge in this line. Young teachers in chemistry frequently continue their collegiate course by attending summer schools, and in this way they increase their knowledge of *pure* chemistry. Such

a practice is extremely commendable and valuable for both the teacher and his school, but unless he makes a special effort outside of the college course to increase his knowledge of chemical applications, his knowledge of the subject and hence his teaching will continue to be one-sided and much less fruitful than if he also increases his knowledge of practical affairs.

The University of Texas Requirements for One Unit Entrance Credit.

The suggestions and directions in this paper need not be followed to secure affiliation in chemistry at the University of Texas. They are offered to teachers as aids in shaping their courses in chemistry, and whether or not these suggestions are followed has nothing to do with securing affiliation. The University aims to be exceedingly liberal and broad-minded in its requirements for affiliation as shown by the following list of requirements.

1. A properly prepared teacher:—see page 9.
2. A proper class-time allowance. The time spent in the course *must not* be less than three recitation periods of 45 minutes and two 90 minute laboratory periods a week for one year.
3. A fairly well equipped laboratory in which the students do individual work.
4. Some evidence that the (competent) teacher occupies the time with a well planned and well conducted course. This evidence is obtained by a personal visit of an official, and an inspection of the note books and examination papers of the students in the course.

Some Suitable Text Books in Chemistry.

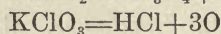
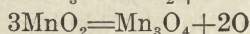
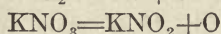
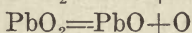
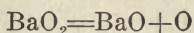
- Brownlee & Others, *First Principles of Chemistry*, Allyn & Bacon.
- Hollis Godfrey, *Elementary Chemistry*, Longmans, Green & Co.
- Hessler and Smith, *Essentials of Chemistry*, Benjamin Sanborn Co.

- McPherson and Henderson, *Elementary Chemistry*, Ginn & Co.
- Morgan & Lyman, *Elementary Text on Chemistry*, MacMillan Co.
- Newell, *Inorganic Chemistry*, D. C. Heath & Co.
- B. W. Peet, *Laboratory Manual in Chemistry*, published by the author at Ypsilanti, Mich.
- Remsen, *Introduction to Chemistry* (Briefer Course), H. Holt & Co.
- Smith, *Laboratory Outline of General Chemistry* (4th Edition) Century Co. This is the book referred to in Part III where *Smith* appears. It is primarily a college text, but is likely to be in the hands of every teacher, and the author preferred to refer to it *alone* rather than to *all* the high school texts.
- Smith & Hall, *The Teaching of Chemistry and Physics in Secondary Schools*, Longmans, Green & Co.
- Alexander Smith, *Inorganic Chemistry*, Century Co.
- Julius Stieglitz, *Qualitative Analysis*, Century Co.
- Prescott & Johnson, *Qualitative Analysis*, Van Nostrand Co.
- Roscoe & Schorlemmer, *Inorganic Chemistry*, 2 volumes, MacMillan Co.

PART III.*
OUTLINE OF AN INTRODUCTION TO THE FIRST PRINCIPLES OF CHEMISTRY.†

Oxygen.

1. Demonstration of its preparation by the decomposition of oxygen compounds at high temperatures (*Smith*, Art. 10). The weight relations of the substances which decompose giving oxygen are as follows:



The names of these substances and the above information concerning the quantities involved in the reactions should be committed to memory by the students. "These are *particular* substances which show the *general* behavior of giving up oxygen readily at high temperatures." Note the significance of coefficients and subscripts in chemical formulae and equations. Fundamentally the symbols represent a certain *number of parts by weight* of a particular substance, which *number* is given in an accompanying table commonly referred to as a *table of atomic weights*. The student should not stumble at the word "atomic weights,"—he should simply accept this for the present as merely a name for the table. The *quantitative* significance of the symbols and equations should be emphasized by problems based on them.

2. *Catalytic Effect of Manganese Dioxide in the Decomposition of Potassium Chlorate* (*Smith*, Art. 11).

Avoid presenting MnO_2 in the equation for the decomposition of potassium chlorate; equations are expressions of *quantitative relations only*, and the manganese dioxide undergoes no

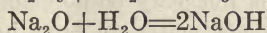
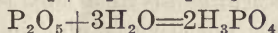
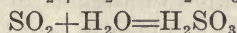
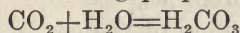
*This part has been issued separately by the author.

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change and is not necessarily present in any particular proportion, hence its quantity is not involved and should not appear in the equation.

3. *Preparation and Collection of Several Bottles Full of the Gas, and the Combustion of Sulphur, Phosphorus, Carbon, Sodium and Iron In It.* Perform this in the usual manner. This shows experimentally that oxygen is very slightly soluble in water, that it is a colorless inodorous gas, that it is as heavy as—not heavier than—air, etc. Students should *learn* these properties *from experiments* rather than from the printed page.

The products of combustion should be treated with water and the resulting solutions tested with litmus. These substances react with water in the following proportions:



Fe_2O_3 does not react with water.

These equations are to be memorized and their quantitative significance to be emphasized with numerical problems. Such reactions with water are called *hydration* and this is a very common chemical phenomenon. Note the *general fact* here illustrated, that *non-metals form acid oxides* and *metals form basic oxides*. This is a fundamental fact and should be remembered.

Hydrogen.

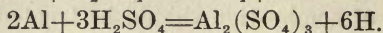
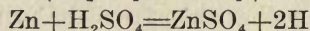
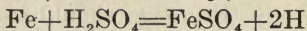
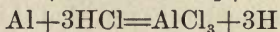
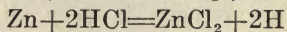
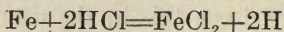
1. General facts concerning its preparation: the interaction of metals and acids:

(a) Many non-oxidizing acids (hydrochloric, dilute sulphuric, acetic, phosphoric) react with any metal above hydrogen in the electromotive force table (e. g. iron, zinc, aluminium, magnesium—see page 110) to form hydrogen and one other product (salt); while there is no action with the metals below hydrogen in the table.

(b) Oxidizing acids (nitric acid, concentrated sulphuric acid) act on almost all metals except the lowest in the table (gold, platinum); but hydrogen does *not* appear as one of the products, and the action is altogether more complex than any action with non-oxidizing acids.

The above general facts (*a* and *b*) should be learned and retained by the student. He should realize that by this means he learns *a great number of important facts easily*.

In connection with (*a*) the following quantitative data is given:



These equations should be committed to memory and drilled on by means of numerical problems.

2. Preparation and collection of the gas to show some of its common properties. Perform as given in any text. The experiment *shows* that the gas is not very soluble in water, that it is lighter than air, that it forms water on burning and that when mixed with air in certain proportions it forms explosive mixtures.

3. Qualitative reduction of some metal oxide (lead oxide or copper oxide), Smith, Art. 20. Point out how this experimental procedure may be used (as it *was used*, by Dumas) to get the proportion by weight in which hydrogen and oxygen combine to form water.

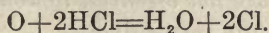
Incidentally the relative rate of diffusion of gases may be demonstrated, as usual, by means of hydrogen and air.

The change of gas volumes through change of temperature and pressure may also be considered here.

Chlorine.

1. The reactions involved in the preparation of chlorine are all complex reactions of oxidation and reduction, hence the preparation of chlorine does not fit into our general plan at this point. However, it is desirable to demonstrate here the properties of the element. For the reason just given, the details of the reactions involved in the preparation of chlorine *are not considered* and the general fact alone is given, namely, that it is prepared by the action of oxidizing agents upon hydrochloric acid.

The oxidizing agents serve in a sense as suppliers of oxygen merely, and the oxygen that they supply combines with the hydrogen of the hydrochloric acid to form water. The following equation represents this change and is the only one that should be given and learned in this connection:



Use several oxidizing agents to illustrate this general method of producing chlorine (Smith, Art. 29a).

2. Prepare and collect several bottles full of the gas to demonstrate some of its properties. Proceed as given in any text.

Hydrogen Chloride.

1. A general method for producing this substance is the reaction between a metal chloride and a *relatively* non-volatile and non-oxidizing (toward HCl) acid. The mixture must not contain water since otherwise the gas dissolves in it. See Smith, Art. 31 (a) and (b). Call attention to the fact that concentrated sulphuric acid is not an oxidizing agent toward HCl while it is towards metals.

The study of *water* as a solvent, as water of crystallization, etc., and of *air*, its composition, etc., may be introduced anywhere here at the option of the instructor.

Acides, Bases and Salts.

A thorough treatment of this topic is advisable at this point. The fundamental relation is "an acid with a base gives a salt and water." This is the *only real criterion* for the definition of the terms *acid*, *base* and *salt*. Such properties as the effect upon litmus paper, taste, etc., characterize only those particular acids or bases which are *soluble* and they are not general or *defining* characteristics. In order that the above relation, that an acid with base gives a salt and water, may really serve to identify a particular substance, we must admit that there are some substances such as sodium hydroxide, calcium hydroxide, etc., which *everybody* agrees to call *bases*, and when one of these reacts with an unknown substance in a manner similar to its reaction with a well recognized acid, then the unknown substance

is *thus recognized* to be an acid; and the corresponding argument is applied when an unknown substance reacts with a substance which everybody agrees to call an acid, such as hydrochloric or sulphuric acid. All other descriptive properties concerning these substances are *not essential to define* or characterize them as acids, bases or salts respectively.

In the experimental demonstration give first the neutralization of a soluble base by a soluble acid with the aid of an indicator; and second the neutralization of a soluble acid by an insoluble base. For instance, add copper oxide to warm dilute sulphuric acid until some of the insoluble oxide remains unacted upon, then concentrate the solution. The salt will crystallize out on cooling.

All acids, bases and salts are *binary compounds*. This should be accepted as a plain fact, without any discussion. The two kinds of constituent parts are called ions, for reasons given later on.

At this point it is advisable to enlarge very much upon the usual treatment of *Acids, Bases and Salts* in text books. For this purpose students should memorize the formulae of about eight to ten basic oxides (e. g. Na_2O , K_2O , CaO , CuO , MgO , ZnO , Al_2O_3 , Fe_2O_3), as well as the formulae for a number of acids in addition to those already met with before (e. g. HNO_3 , H_3PO_4 , H_2CO_3 , H_2SO_4 , H_2SO_3 , $\text{H}(\text{C}_2\text{H}_3\text{O}_3)$).

Next in order, the student should learn to *recognize* the valence of the *metal* ions or of the acid ions *from the formulae* of the bases or acids just given. Without any reference to the atomic theory, define valence as follows: "The valence of an ion (that is, one of the two kinds of parts into which any acid, base, or salt primarily divides) is the number of H's it appears to take the place of, or to combine with, as shown by the formulae of different compounds on comparison."

Note that this definition presents valence as a *numerical relation shown by the formulae* and does not mention the word *atom*. Reference to the atomic theory is thus seen not to be necessary; though there is no objection to considering valence as a property of atoms or groups of atoms if the connection between the formulae and the atomic theory of matter has been pointed out before this.

Next in order, the student should learn how the metal ions and acid ions from the above memorized formulae must be put together in accordance with the demands of valence in order to obtain the formulae of the normal salts. This should be drilled upon, and the fundamentals of nomenclature should be given.

Finally drill the student in writing equations between all the acids and all the bases which have been memorized. They are all metathetical reactions.

Definition of the metathetical reaction: It is one in which two binary compounds exchange their negative parts (or their positive parts) producing only *two* new compounds, the valences of all ions remaining constant during this change. Whenever two substances are said to react *metathetically*, this one word tells *fully what takes place*: with this knowledge, anyone is able to write the equation of a change if he knows the formulae of the substances involved.

Hydration of Oxides.

1. To be illustrated with the slaking of quick lime; then follows the general information:

(a) Basic oxides when hydrated usually appear as hydrated to *the maximum extent*, that is, two hydroxyls are formed from every *O* in the formula of the oxide, e. g., $\text{CaO} + \text{H}_2\text{O} = \text{Ca}(\text{OH})_2$ (*drill in writing such equations with other basic oxides*).

(b) When placed in contact with water, only those oxides the hydroxides of which are soluble actually take up water to form hydroxides. Of the common hydroxides only those of the alkali and alkaline earth metals are soluble. Drill!

2. The *hydroxides* of metals also are *bases*, just as well as the oxides. Drill on writing all the equations of the reactions between the hydroxides of the eight metals (in the list of bases memorized above) with the acids given above.

3. Recall the fact that metals form basic oxides and non-metals form acid oxides. The *latter* hydrate just as well as the basic oxides; however, the extent of hydration follows *no rule*, and is *usually less* than the maximum. Common acid oxides are: SO_2 , SO_3 , CO_2 , P_2O_5 , N_2O_5 , N_2O_3 , P_2O_3 . Drill on writing

the relation between these plus water, giving the ordinary acids, (H_2SO_3 , H_2SO_4 , H_2CO_3 , H_3PO_4 , HNO_3 , HNO_2 , H_3PO_3).

4. The actual hydration of acid oxides on contact takes place readily with nearly all acid oxides because the resulting acid is soluble in nearly all cases (exceptions: silicic acid H_2SiO_3 , from SiO_2 ; arsenious acid, HAsO_2 , from As_2O_3).

Solubility of Salts.

By solubility is meant the solubility of *these substances in water*, not their solubility in a solution which acts on them chemically. None of these substances are absolutely insoluble; they are all slightly soluble and they present *enormous relative differences* in their actual solubilities.

A short table, of the solubilities of the ordinary salts such as is given below, should be committed to memory by the student.

Table of Solubilities of Salts.

Nitrates and Acetates. All soluble.

Sulphates. All soluble except PbSO_4 , BaSO_4 , SrSO_4 , and CaSO_4 . The last is perceptibly soluble.

Chlorides. All soluble except AgCl , HgCl , PbCl_2 . The last is slightly soluble in cold water, and quite soluble in hot water.

Normal Carbonates and Phosphates. All insoluble except those of the alkali metals, Na, K and " NH_4 ."

The solubilities of the substances is one of the largest factors which determine whether or not metathetical reaction will take place. Thus even in the reaction between a base and an acid, for which there is always a decided tendency, insolubility will retard the action somewhat in proportion to the insolubility of the salt formed, and may prevent it practically entirely. The following experimental procedure serves to show this. The experiment is an excellent one to develop *chemical notions*:

Put into each of three test-tubes a pinch of zinc oxide, and add to one just enough dilute HCl so that when the mixture is stirred and warmed the zinc oxide dissolves. Treat the second

portion of zinc oxide with HNO_3 , and the third with dilute H_2SO_4 . In the same way try calcium oxide (or hydroxide), and lead oxide. Do you observe any relation between the solubilities of the salts that are, or should be, formed, and the rate at which reaction takes place?

When insoluble salts are, or would be, thus formed, is there absolutely no action or is the action merely retarded? To find the answer to this question examine particularly the mixtures of lead oxide with HCl and H_2SO_4 , respectively. The change from yellow oxide to white salt in these cases will reveal whether or not any action takes place. If you conclude that the action is merely retarded, and you wish to find an explanation of this, drop a small lump of marble into some dilute H_2SO_4 and observe that a copious evolution of CO_2 takes place for just a moment, followed by a very slow action. Take out the piece of marble, scrape off the outer layer and drop the lump into dilute H_2SO_4 again. Rapid action which lasts only for a short while is again observed. It is evident that the layer of insoluble calcium sulphate formed covers the lump and retards the diffusion of the acid to the calcium carbonate. In the experiments above, the particles are much smaller, yet the actions are retarded in the same way.

Acid Salts.

With due consideration of what the ionisation theory teaches concerning the ionisation of polybasic acids, *acid salts* should be treated somewhat as follows: when an acid such as phosphoric, H_3PO_4 , separates into its main constituent parts (ions), at first only one H separates, leaving the remainder intact as a monovalent ion $(\text{H}_2\text{PO}_4)'$; and when the first portions of a base are added to a solution of such an acid, the reaction of neutralization takes place as though this were a monovalent acid, the composition of the acid radical of which is $(\text{H}_2\text{PO}_4)'$. Hence the aluminium salt would have the formula $\text{Al}(\text{H}_2\text{PO}_4)_3$.

Note:—the valence of *metals* and all positive ions is designated with an asterisk—e. g. H^* ; that of acid radicals and other negative ions by an apostrophe, as NO_3' , SO_4'' , etc.

After enough base has been added to neutralize all of the first

set of H-ions, then the second H ionizes extensively and leaves the bivalent ion $(\text{HPO}_4)''$; hence salts formed under these conditions appear to have a formula similar to those formed with bivalent acid ions, e. g., $\text{Al}_2(\text{HPO}_4)_3$.

After the second hydrogen has been neutralized, then the tendency to ionize the third hydrogen may become effective, leaving the trivalent ion, $(\text{PO}_4)'''$.

Of course if enough base is added all at once to neutralize more than one or two *H*'s the reaction immediately proceeds to the corresponding extent. Thus if a drop of sulphuric acid is added to more than enough of sodium hydroxide solution to neutralize the acid completely, then the salt Na_2SO_4 is formed immediately; but when the procedure is reversed, that is, when a drop of sodium hydroxide solution is added to a great deal of sulphuric acid, then it will form only the acid salt $\text{Na}(\text{HSO}_4)$.

That the (experimental) formation of acid salts, with polybasic acids, *depends only on the relative proportions* of acid and bases *mixed* should be strongly emphasized and drilled upon. In the following lessons the reactions between carbon dioxide and lime water; and the passing of SO_2 into sodium hydroxide solution until no more is absorbed—these and others furnish excellent examples to bring out this same point; and this reaction also furnishes a good opportunity to develop chemical notions as distinct from the mere arithmetic notions involved in equation writing.

The following experiment is here given in full to demonstrate the actual formation of acid salts:

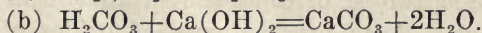
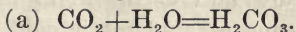
“Secure a fresh solution of tartaric acid, $\text{H}_2(\text{C}_4\text{H}_4\text{O}_6)$, containing about 200 grams per liter, and a solution of potassium hydroxide containing about 150 grams to the liter and fill two burettes with these solutions respectively. Measure out 20 cc. of potassium hydroxide solution into a small flask, add 30 cc. of the tartaric acid solution measured from the other burette, then heat the mixture to the boiling point, add a drop or two of phenolphthalein solution and finish neutralizing it by adding more tartaric acid from its burette. Cool the mixture under a jet of tap water and add to it as much tartaric acid again as was necessary to neutralize the potassium hydroxide. Crystals of potassium acid-tartrate will be formed. Next heat the mixture

to boiling and add KOH slowly, with constant stirring until an amount approximately equal to the original amount has been added. Note that the crystals dissolve as KOH is added. Explain what happens in each step and write the equations of the three reactions; they are all metathetical changes. Note that the second portion of potassium hydroxide used is equal to the first portion used. The composition of the crystals that separated from the solution is $\text{KH}(\text{C}_4\text{H}_4\text{O}_6)$, potassium acid-tartrate, or potassium bitartrate. Give reasons for both names.

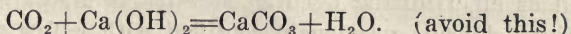
Carbon.

Present some of the properties of the element, the discussion of the flame, some descriptive matter concerning the importance of carbon monoxide on account of its presence in water gas and its use in the reduction of ores (e. g. of iron ores).

2. Show that the preparation of carbon dioxide from practically any carbonate is in general similar to the preparation of HCl (noting however the difference in the solubilities of the two gases and that carbon dioxide is not oxidizable; hence almost any ordinary acid may be used to liberate it from its salts). Show some of its common properties and emphasize that the portion dissolved in water is largely hydrated; whenever it acts, in solution, as an acid on a base, then it is H_2CO_3 , which is the actually acting substance, and not CO_2 . A separate equation should always be written to express the hydration and dehydration that may precede or follow a metathetical reaction. Thus when carbon dioxide is absorbed by lime water the following reactions take place in the order here given:



These reactions should *not* be combined into the following expressions:



Be certain to make the experiment of passing carbon dioxide into lime water until the solution clears up again; and then boil the solution. The student should know this experiment and action *thoroughly*, and also know its application in nature and in the production of boiler scale (temporary hardness of potable water).

Sulphur.

Deal briefly with the properties of the element (the phenomena presented by melted and cooled sulphur, etc., need not be considered). Emphasize the source and extraction of commercial sulphur.

Deal with hydrogen sulphide very briefly here. Make some iron sulphide and liberate hydrogen sulphide from it. Note that the gas is poisonous. The reactions involved in these examples are mainly simple or metathetical, and the complex oxidation reactions of hydrogen sulphide must be omitted here. The use of hydrogen sulphide as a laboratory reagent will be sufficiently illustrated later.

Prepare sulphur dioxide by means of the reaction between sulphites and acids. Point out that the method in general is the same as that employed for the production of carbon dioxide and of HCl. The student should learn this and also remember the particular substances that he actually handled in this connection in the laboratory. In the equations involved, the expression for hydration and dehydration should be separated from the metathetical reactions—as was advised for carbon dioxide. The main chemical property of SO_2 (or of sulphurous acid) is its tendency to be oxidized to sulphuric acid. This may be illustrated as follows:

Prepare a saturated aqueous solution of SO_2 . Put a little of it in a test tube, add a few drops of hydrochloric acid and a drop or two of barium chloride solution. Only a slight or negligible precipitate will be formed. Now add to the main portion of the solution one or two cc. of concentrated nitric acid and heat gently to boiling. Some brown fumes of oxides of nitrogen will be formed to a slight extent, which show that the nitric acid has acted as an oxidizing agent. Now test the solution again with barium chloride previously acidified with a few drops of HCl. A copious precipitate of barium sulphate will be obtained. This test depends upon the fact that barium *sulphite* is soluble in this mixture while barium sulphate is not; hence not until the sulphurous acid is changed to sulphuric will a precipitate be formed.

The commercial production of sulphuric acid by the contact

method may be given here. If the old English method is given, no reference need be made to the formation of nitrosyl sulphuric acid. It is perfectly correct to state that *NO* takes up oxygen to form NO_2 and this change takes place more readily and hence faster than the *direct* reaction between SO_2 and the oxygen of the air. The oxides of nitrogen are merely *catalytic agents*, and the impelling tendency of the action is always the tendency of SO_2 to take up oxygen.

Ammonia.

Stress the commercial source of ammonia; its liberation from the salts by a general method which corresponds to the general method for the liberation of carbon dioxide, HCl , and SO_2 . Any ammonium salt and any base may serve for this purpose although the reaction takes place readily only with a strong (soluble) base. Stress the fact that ammonia hydrates readily—otherwise the usual treatment found in texts may be given.

Other Optional Topics.

The following topics may be treated here at the option of the teacher:

1. The liberation of nitric acid from its salt,—but do not consider at this point the production of oxides of nitrogen, or the reaction of nitric acid with metals, etc.

2. The other halogens: fluorine, bromine and iodine.

For fluorine the etching effect of hydrofluoric acid upon glass is the only experiment to be given. For bromine and iodine, show, with test tube trials, the liberation of HBr and HI from their salts by means of phosphoric acid (do not use sulphuric), and then show the liberation of the elements from these acids by means of an oxidizing agent (add powdered MnO_2 to the mixtures in the test tubes): The displacement, by chlorine, of bromine from bromides and of iodine from iodides, should also be presented here.

*Ionisation and the General Relation Between Dissolved Substances Which Results in Metathetical Reaction.**

This topic should be begun by *showing* experimentally some fundamental facts concerning electrical conductivity of solu-

*Read pages 47-49 in this connection.

tions. This may be shown most conveniently by means of an alternating electric light current and an incandescent lamp. Cut one of the two wires of an "extension cord" and, to the two ends thus obtained, attach platinum wires for electrodes. Dip the latter into a beaker filled with distilled water, attach the plug to the light circuit, and insert the lamp into the socket of the cord. It will be found that distilled water does not conduct the current sufficiently well to light up the lamp. Next pour a little hydrochloric acid into the water: the lamp will burn brightly, which shows that this solution conducts well. Repeat with fresh water and acetic acid: the lamp will burn dimly, which shows that the solution conducts poorly. In the same way try sodium hydroxide solution, ammonium hydroxide solution, alcohol in water, and various salt solutions. These experiments suffice to dispose the student favorably to receive the following general facts.

Acids, bases, and salts, when dissolved in water dissociate into parts, called ions, each of which carries a definite number of positive or negative unit charges, the number of which corresponds to the valence. Metals and hydrogen carry positive charges, while the non-metals, acid radicals, and the hydroxyl radical carry negative charges.

Not all of an acid, base, or salt in a solution is present in the form of ions. A part is present in the undissociated or combined form. This is due to the tendency of ions to recombine. These two *opposing* tendencies, of ionisation and of recombination, hold each other in equilibrium *when* a certain fraction of the dissolved substance is in the form of ions and undissociated part, respectively. These fractions have different values with different substances, and the "ion" fractions increase with dilution, while the *undissociated* decrease.

The following general statement covers the degrees of dissociation of all common substances. It should be committed to memory:

All normal salts, all strong acids (hydrochloric, nitric, sulphuric), all strong bases (that is, all *soluble* bases except ammonia) ionize extensively in aqueous solutions.

All weak acids (acetic, carbonic, sulphurous, phosphoric), all weak bases (ammonia) and water itself, are very slightly dissociated (about 1% and much less) in aqueous solutions.

Insoluble or very slightly soluble substances can not form many ions in solution and hence should be included here as slightly dissociated substances.

The chemical activity of acids, bases, and salts depends upon the concentration of their ions; in other words, it is primarily the ions, rather than the undissociated portions, which take part in chemical reactions. This is suitably illustrated by means of the rates of action of hydrochloric acid upon zinc, and of acetic acid upon zinc, or the rates of action of these acids upon marble, all of which is readily shown by means of test-tube trials.

The fact that a dissolved acid, base, or salt is present in solution in two forms, either one of which tends to change to the other one until they hold each other in equilibrium,—this fact appears strange to the beginner, and yet it is similar to the familiar phenomenon of a salt dissolving in water until the solution is saturated, and hence it is in equilibrium with the solid salt; it is also similar to the evaporation of a liquid until its saturated vapor is in equilibrium with it. However, there is one essential difference between these equilibria and the equilibrium between a salt and its ions, namely, there are always two kinds of ions, and neither one kind alone can hold the equilibrium with the undissociated salt. *Both* must be present, but their amounts or numbers per cc. necessary for any one state of equilibrium need not be equal,—only the product of these numbers per cc. must remain the same. Thus, if 10 H ions and 10 acetate ions (per cc.) are present in a certain solution (and hence holding in equilibrium the undissociated acetic acid present) then this equilibrium would also be held by 1 H ion with 100 acetate ions (99 of them from another acetate), or by 1 acetate ion with 100 H ions (99 from another acid), *because* in all three instances the product would amount to 100. In other words, the effect of the ions in the equilibrium relation with their compound is *measured* by the *ion-product*.*

The following experiment, which involves exactly the same relation, demonstrates its operation in the simplest manner.

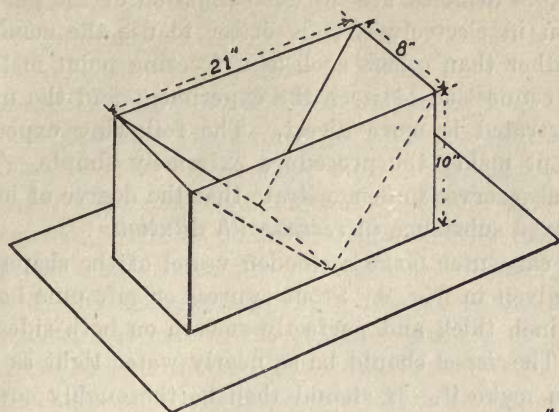
Prepare one-half test tube full of a cold saturated solution

*This relation does not hold *exactly* with greatly dissociated acids, bases, or salts; but even with these, the *fundamental* relation is probably the same, and hence the illustration serves for all.

of naphthalene in absolute alcohol and one-half test tube full of a cold saturated solution of picric acid. Pour into a small (clean and dry) flask one-half of the clear saturated solution of picric acid and two-thirds as much of the naphthalene solution, the two quantities being gauged as accurately as possible with the eye. They unite to form naphthalene picrate, designated by *NP*. This solution and all others in this experiment must be kept cool,—they should not be warmer than 18 degrees C. Allow the mother liquid to drain from these crystals, and prepare a saturated solution of *NP* by adding a small portion of absolute alcohol, shaking the mixture vigorously, and repeating this treatment until most of the crystals (not all!) have dissolved. To one-half of this solution add one-eighth as much (by volume) of the remaining clear picric acid solution. Some of the compound *NP* should separate from the solution. To the other half of the saturated *NP* solution add one-eighth as much of the clear naphthalene solution. Again some of the compound *NP* should separate from the solution.

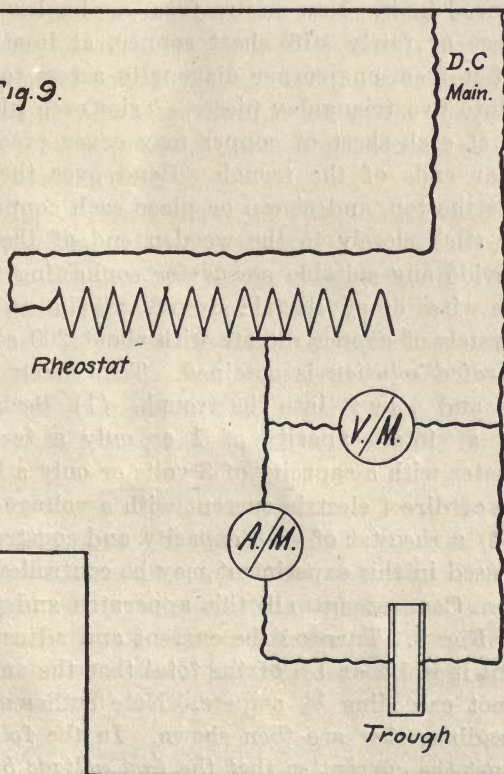
Since in one test-tube full of solution of *N P*, the addition of *N* resulted in the formation of more of the compound, there must have been some (free) picric acid in this solution; and since in the other test-tube full of this solution the addition of *P* resulted also in the formation of more of the compound, it follows that the solution also contained some (free) naphthalene. Hence the solution contains both free *N* and free *P* besides the compound *NP*. The effect of the free *N* and free *P* upon the equilibrium with their compound *NP* is measured by the product of their numbers per cc., say $a \times b$ (if a is the number of *N* per cc. and b is the number of *P*). When a is increased by adding a little of the saturated solution of naphthalene, then $a \times b$ is increased and the equilibrium disturbed. In order to regain equilibrium, some of the free *N* and free *P* combine, thus reducing a and b until $a \times b$ reaches its former value again; and this newly formed *NP* separates from the solution because the latter is *saturated* with *NP*. The corresponding change takes place when picric acid is added.

Fig 8



Dimensions in Inches — "Inside!"

Fig. 9.



Conductivity Trough and Connections

In order to demonstrate the determination of the per cent of dissociation in electrolytes it is better to use the conductivity method rather than others such as a freezing point method, because the connection between the experiment and the notion to be demonstrated is more direct. The following experimental arrangement makes the procedure extremely simple. This experiment also serves to demonstrate that the degree of ionisation of a dissolved substance *increases with dilution*.

Have a carpenter make a wooden vessel of the shape and dimensions given in Fig. 8. Stout cypress or soft pine boards, at least $1\frac{1}{4}$ inch thick and perfectly smooth on both sides, are to be used. The vessel should be as nearly water tight as the carpenter can make it. It should then be thoroughly covered on the inside with melted paraffin to make it absolutely water tight, and in order that the boards may not take up any of the solutions poured in it. Now secure from a plumber or cornice maker a piece of fairly stiff sheet copper, at least 16 inches square. Cut it from one corner diagonally across to the opposite corner into two triangular pieces. Trim each piece if necessary, so that each sheet of copper may cover exactly one of the triangular ends of the trough. Bend over the excess of each plate at the top, and clamp or place each copper plate so that it may stick closely to the wooden end of the trough it covers. Provide any suitable means for connecting the copper sheets to the wires of an electric circuit. Shake up about 200 grams of crystals of copper nitrate with about 200 cc. of water until a *saturated solution* is obtained. Take about 100 cc. of this solution and pour it into the trough. (1) Secure an ammeter with a total capacity of 1 or only a few amperes, (2) a volt-meter with a capacity of 3 volts or only a little more, (3) a source of direct electric current with a voltage of 2 to 10 volts, and (4) a rheostat of such capacity and construction that the current used in this experiment may be controlled to within 0.01 ampere. Connect up all this apparatus and the trough as shown in Fig. 9. Turn on the current and adjust it with a rheostat until it is 1.5 or 1.6 of the total that the ammeter can carry, but not exceeding $\frac{1}{2}$ ampere. Note both ammeter and voltmeter readings that are then shown. In the following operations adjust the current so that the *first voltage between the*

poles (which the voltmeter indicated) *is kept constant*, and record in parallel columns the amount of current that flows *after each* dilution of the solution. Dilute the solution by adding measured amounts of distilled water: at first in portions of 100 cc. at a time, subsequently in larger portions of several 100 cc. At the beginning, the increase in current will be relatively large with each addition of water; then it will become less until finally no further increase in current is obtained.

Divide the final (maximum) value into the first value (obtained with the original solution): this gives the fraction of the salt present as ions in the original solution.

The larger current obtained after dilution shows that the number of ions arriving at the poles *each second* is larger after dilution than before. Since the attractive force (voltage) is kept constant and hence the ions move at the same speed, and since on dilution they are moved parallel to the poles but remain, as a whole, at the same perpendicular distances from them, there remains no other way to account for the greater number of ions arriving at the poles after dilution except the conclusion that the number of ions is increased with dilution until all the ions possible have been formed.

Point out next the general conditions under which metathetical reactions take place, and apply this to the neutralization of an acid with a base and to the experiments given below. The following outline may be helpful here.

(a) Whenever two solutions (or a solution and a slightly soluble solid) are mixed, the (accidental) meeting of the cations from one solution with the anions from the other solution will produce at least small amounts of all the new compounds possible.

(b) Note the amount (in a general way, that is, whether large or small) of each free ion in the mixture at the beginning.

(c) Note the amount (i. e. large or small) of free ions that can be produced by each of the resulting substances. This is the amount of its ions with which each one would be in equilibrium,—this is very small in the case of slightly soluble and slightly dissociated substances.*

*To be exact, we should consider here the product $a \times b$, of the concentrations of each pair of ions in place of their *amounts* (see page 75).

(d) If an insoluble or slightly dissociated substance is among the resulting ones, and the amount of its ions in the original mixture is much larger than the small amount with which it would be in equilibrium,* then this pair of free ions will combine and thus they will reduce their amounts until these are small enough for equilibrium. As thus the free ions disappear, any undissociated portions of their original compounds will ionize and be used up in turn.

If one of the original substances is only slightly soluble (as in *d* below), then at first it will dissolve only in the small amount which saturates the solution. Then as the amount in solution is changed by reaction, more solid will dissolve, and so on.

All metathetical reactions are due to such extensive reduction of the original number of a certain pair of ions. A *slight* formation of a new compound, by ions combining to a slight extent, is not considered to be a reaction.

The student should make test-tube trials with the following mixtures, and point out in each case the particular pair of ions which by *combining extensively* serve in a sense as the *primary* cause of the reaction:

(a) Mix any one of several barium salt solutions with any one of several sulphates, producing in this way barium sulphate from at least nine different mixtures.

(b) Produce silver chloride from several different mixtures.

(c) Produce ferric hydroxide from several different mixtures.

(d) Dissolve calcium phosphate in dilute hydrochloric acid. Here the least ionized combination is a combination of H ions with PO_4 ions, for instance, $(\text{H}_2\text{PO}_4)'$, the calcium salt of which acid radical is soluble.

(e) To the mixture obtained in (d), add ammonia to neutralize the acid. The ammonium phosphate thus produced will introduce many PO_4 ions, and calcium phosphate will be obtained as a precipitate.

(f) Add solution of sodium carbonate to an aqueous solu-

*Or, to be exact, the original amounts of its ions form a larger *ion-product* than that with which the insoluble or slightly-dissociated substance can be in equilibrium.

tion of the salt of any metal that forms insoluble carbonates. Filter and dissolve the precipitate by means of the addition of any acid that forms a soluble salt. Repeat with the salts of five other such metals.

These experiments enable the student to see that whenever he knows the general relations between the ionized substances in a given mixture, then he may predict whether or not reaction will take place.

Exercise.

Which of the substances given below when mixed will react? give a reason for your answer.

1. Lead oxide and water.
2. Barium oxide and water.
3. Calcium carbonate and hydrochloric acid.
4. Sodium acetate solution and hydrochloric acid.
5. Copper sulphate and sodium phosphate solution.
6. Since dry sodium chloride and concentrated phosphoric acid react to form HCl by metathesis, what are likely to be the ionisation relations in this liquid medium (concentrated phosphoric acid) to bring about this change?
7. If the liquid medium in (6) is changed by the addition of much water, will any reaction take place extensively? What are the ionisation relations in this latter case.

Other questions of *similar character may be added.*

Electrolysis.

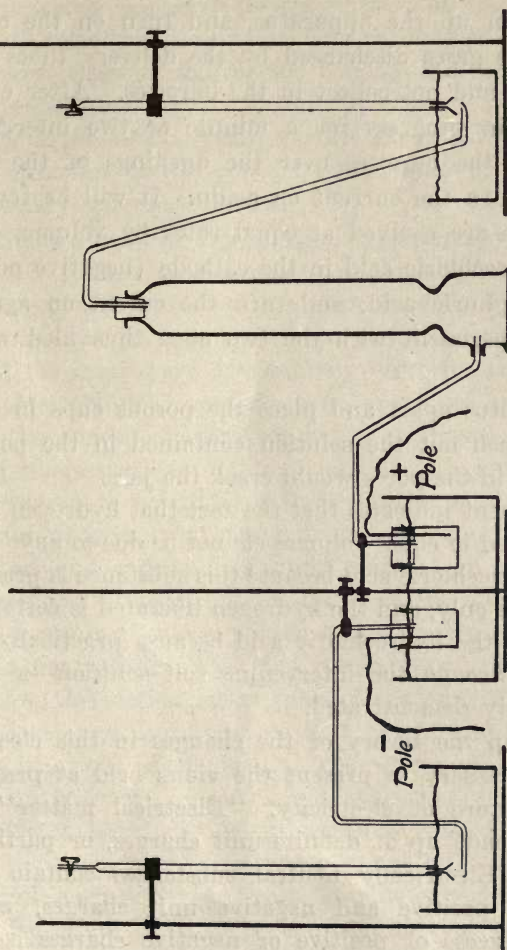
The first experiment to be given is the electrolysis of hydrochloric acid, and this should be carried out with the following apparatus: Secure two small porous cups, 3 to 4 inches high and $1\frac{1}{2}$ to 2 inches in diameter. Dip the upper edges to a depth of one inch into melted paraffin in order to close up the pores in that portion of the cup. Fit rubber stoppers to the

cups. Through each stopper cut two holes,—one to fit a piece of retort carbon, the other to fit the glass tubing with which the apparatus is to be connected.* Secure also a porcelain jar 4 or 5 inches in depth and 5 or 6 inches in diameter in which the porous cups are to be placed: This jar is to be filled with a saturated salt solution. Secure a tall, slender bottle, of at least one quart capacity, fit it with a rubber stopper and two glass tubes one of which extends to the bottom of the bottle while the other terminates just below the stopper. In place of this bottle, the cylinder shown in Fig. 10 may be used. This bottle or cylinder serves to retain the chlorine, and allows an equal volume of air to be discharged in place of the chlorine. Air may be collected over water without appreciable loss, while chlorine cannot be collected over water because it is too soluble. Glass tubes for connections should now be bent as shown in Fig. 10. The tube fitted to the cup on the left should have no rubber joints in it, because hydrogen is to be evolved at this pole, and this gas would diffuse through the rubber tubing joints, and thus vitiate the experiment. The ends of the delivery tubes which are to be placed under the burettes should be drawn out to a small opening. Secure two burettes and two dishes or beakers full of water. Place the burettes in position with the lower ends extending into the water, and fill them by drawing up water by means of a piece of rubber tubing. By this means they may be quickly refilled when the experiment is to be repeated. For electric connection, twist some bare copper wire around two arc light carbon rods, insert these through the rubber stoppers into the porous cups, and connect the copper wire with the terminals of an electric circuit which supplies *direct* current at a voltage of 10 to 25 volts. The electrode on the right should be connected to the positive terminal. Insert a switch in the circuit with which the current may be conveniently turned on or off.

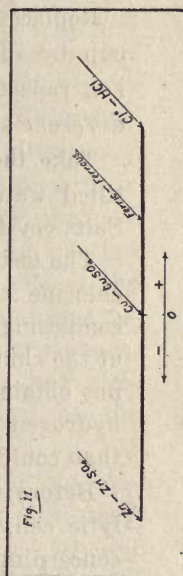
Fill the porous cups three-fourths full of a mixture of equal

*For cutting holes in rubber stoppers, sharpen the edge of a cork borer by means of a file, thus producing a rough saw-like edge which is very effective, dip the end of the borer into caustic soda solution or into alcohol and proceed as with cork stoppers.

Fig 10



Electrolysis of Hydrochloric Acid.



Plot of Pole-Voltages.

parts of concentrated hydrochloric acid and water. To the cup on the right add a few crystals of potassium permanganate. By this means the liquid is immediately saturated with chlorine. Now join up the apparatus, and turn on the current, but allow the gases discharged by the delivery tubes to escape into the air and not collect in the burettes. After electrolysis has been in progress for a minute or two interrupt the current, place the burettes over the openings of the delivery tubes and turn the current on again. It will be found that the two gases are evolved at equal rates by volume.

Replace the hydrochloric acid in the cathode (negative pole) cup by dilute sulphuric acid, and turn the current on again, i.e., repeat the experiment with the two cups thus filled with *different* solutions.

Take the apparatus apart and place the porous cups in distilled water to leach out the solution contained in the pores. Salts crystallizing in the pores would crack the jars.

The last experiment indicates that the fact that hydrogen and chlorine are obtained in equal volumes cannot be due to any "decomposing" of hydrochloric acid because this substance is present at the chlorine pole only, and the hydrogen liberated is certainly not obtained from the hydrochloric acid because practically no hydrogen passes *through* the intervening salt solution—a fact that could be easily demonstrated.

Before taking up the theory of the changes in this electrolytic cell, it is necessary to present the views held at present concerning the nature of electricity. "Electrical matter" is considered to be made up of definite unit charges, or particles of definite size. Electrically neutral substances contain an equal number of positive and negative unit charges, while ions contain an excess of positive or negative charges corresponding to their valence.

Only negative charges are mobile, and may move from one material particle to another, or from one kind of matter to another. These mobile negative charges are called *electrons*. The positive charges are fixed upon each particular particle of matter, and are never transferred from one particle to another.

In accordance with this theory, the change at the cathode takes place as follows: the electromotive force of the battery forces electrons to pass from the cathode to those hydrogen ions which are next to the surface of the cathode (negative pole) thus changing these ions to *neutral* hydrogen,—i. e. ordinary, gaseous hydrogen. During the same period of time, an equal number of chlorine ions which are next to the surface of the anode (positive pole) give up their negative charges or electrons to the pole, thus changing to neutral chlorine or ordinary gaseous chlorine.

Since a drop of any solution is electrically neutral to the outside it follows that it must contain *as many* ions with positive charges as it contains ions with negative charges. When ions are discharged at the cathode, the drops of liquid which have lost the cations are momentarily left with an excess of negative charge while the opposite state of affairs exists simultaneously at the anode. These drops at the extreme end act upon their neighbors so as to obtain from them the kinds of ions that are necessary to re-establish their electric neutrality. This action communicates itself from drop to drop, from one pole to the other, and results in a slight shift of all the positive ions toward the cathode and all the negative ions toward the anode, as a result of which all the drops of solution regain their electric neutrality. With the discharge of the next ions at the poles, this action again takes place and so on.

These are the essentials that have to be brought out in connection with electrolysis. The writer realizes that this treatment is very dogmatic, and that there has been no attempt made to show why scientists have arrived at the fundamental notions which are involved here, but he believes that the young mind is scarcely ready to follow the experimental evidence.

However, it must be emphasized that the fundamental notions presented above are very important, and should be taught in the manner in which they are presented here.

The following further experimental demonstrations of electrolysis should now be brought before the students.

Make a 10 or 15 per cent solution of zinc chloride, and fill two porous cups with this solution. As before, place the two cups in a jar filled with salt water, and into one place

a carbon rod for an anode, and into the other a strip of sheet copper for a cathode, but leave the cups open—i. e. unstoppered. Make the necessary electrical connections and pass a current through this cell. Zinc will be deposited at the cathode, and will show itself by its gray color; while chlorine will be evolved at the anode and will reveal itself by its odor.

Replace the zinc chloride in the cathode (negative pole) cup by a copper sulphate solution, insert a clean sheet copper pole and turn on the current again. Copper will be deposited upon the cathode, while the reaction at the anode is the same as before.

Show next the electrolysis of dilute sulphuric acid, using either the apparatus used above for the electrolysis of hydrochloric acid, or a Hoffmann apparatus (see page 53 with reference to the purchase of the latter.)

Next fill the same apparatus with a solution of sodium sulphate in place of sulphuric acid and show that when this solution is electrolyzed, hydrogen and oxygen are liberated in the same proportion by volume as with sulphuric acid. Do not attempt to force a large current through this solution because it is a poor conductor, and the heat effect of the current is liable to crack the apparatus at the electrodes.

Then secure two clean porous cups, the pores of which contain neither an acid nor an alkali. Fill them with a solution of sodium sulphate, insert platinum electrodes, and place the cups in a salt water jar as before. Before the current is turned on, show with the aid of litmus paper that the solution in the cups contains neither an acid nor an alkali. Then turn on the current, and show that the solution in the cathode cup becomes alkaline and the solution in the anode cup becomes acid as the result of the passing of the current.

Replace the liquid in the cathode cup by a dilute solution of copper sulphate, and show that copper is deposited when the current is passed through this apparatus.

These experiments show that the nature of a chemical change at a pole during electrolysis is determined by the nature of the substances present right at the pole, and that it is not dependent upon the nature of the substances present at the other pole or anywhere else.

The electrolytic discharge of the ions *zinc, copper, hydrogen,*

and *chlorine*, requires no further explanation; but the discharge of hydrogen out of the solution of a sodium salt and of oxygen out of a solution of a sulphate require further consideration. For an understanding of these phenomena, the following fundamental or general facts must be recognized:

1. Water and all aqueous solutions contain the ions of water (i. e. hydrogen and oxygen ions practically), and although the amounts of these ions present are small, yet if they are ever used up, more will be formed as fast as the others are used up.

2. If two or more different ions which may be discharged are present at a pole, that particular one will be discharged which will set up the least opposing electromotive force after its discharge. In this connection see page 91, and also page 112. In other words, the ion discharged is the one requiring the least voltage for its discharge.

With these two fundamental facts the discharge of hydrogen from a solution of a sodium salt is readily understood. Since the voltage required for the discharge of the sodium ions is much greater than that required to discharge the hydrogen ions which are also present, the latter are discharged; and the hydroxyl ions which are formed from the water as the hydrogen ions are discharged impart the alkalinity to the solution, or in other words, they together with the sodium ions form sodium hydroxide.

Since oxygen is obtained during electrolysis at a platinum anode surrounded by a solution of a sulphate, it follows that the oxygen ions of the water are more easily discharged than the sulphate ions themselves; and the hydrogen ions which are formed from the water as the oxygen ions are discharged impart the acidity to the solution, or in other words, they together with the sulphate ions form sulphuric acid.

The Electro-Motive Force of Battery Cells.

The subject of electrolysis has not been presented fully unless the electro-motive force of the product formed at the poles has been pointed out, and since this electro-motive force when considered by itself presents the cell as a galvanic battery, it is advisable to consider the latter subject in this connection.

In the treatment of this topic as in the treatment of elec-

trollysis, it must be realized that the fundamental fact is the utter independence of the chemical actions which take place or tend to take place at the two poles. Without this a logical presentation of the facts involved becomes practically impossible and it is on this account that the methods of presentation now ordinarily employed are comparatively valueless, if not actually harmful.

For the following experimental demonstration, secure four small porous cups, a jar with strong sodium chloride solution, and a sensitive voltmeter with a total range of three volts or very little more. Fit the cups with cork or rubber stoppers, which are to be perforated to fit the electrode rods mentioned below. With one porous cup make a chlorine pole by using a carbon electrode rod, filling the cup with a mixture of equal parts of concentrated hydrochloric acid and water, and then saturating the solution with chlorine (preferably by electrolysis, otherwise by adding a little potassium permanganate). For the second pole, fill the cup with zinc sulphate solution, and use a rod of zinc for the electrode rod. For the third pole, use a platinum pole (see page 33)—or, less suitably, a carbon rod, and fill the cup with ferric chloride solution to which has been added a little of a ferrous salt (FeSO_4). For the fourth pole use a copper rod and copper sulphate solution.

Put the copper sulphate pole into the salt solution jar and then put with it in turn each one of the other three poles and measure the voltage between each of them and the copper sulphate pole. Now plot the results on a line as follows: mark a point on the line to denote the voltage of the copper-copper sulphate pole and refer to it arbitrarily as a "zero voltage." Since the zinc sulphate pole is the negative pole when combined into a cell with this copper-copper sulphate pole, and since the combination has a voltage of (about) 1.1 volts, then if we consider the force of the copper pole to be zero (arbitrarily), it follows that the voltage of the zinc pole is -1.1 volts. To represent this on our plot we measure from the copper pole point eleven spaces of any arbitrary length to the left and mark this point (see Fig. 11). Then lay off the voltage of the chlorine pole (which is about $+1.0$ volt) and the voltage of the ferrous-ferric salt pole (the lat-

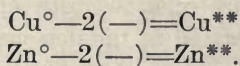
ter will be about $+.4$ to $+.5$ volts.) Measure next the voltage of all other combinations of any two of these poles and compare the values obtained with the number of unit lengths between them shown in the plot. It will be found that, allowing for slight inaccuracies due to this rough procedure, the numbers will agree. This shows that the voltage of a cell is the sum of the two independent voltages of the poles.

The *primary cause* of the action of a pole is the natural tendency to give up or take up electrons which some substance present at the pole possesses; e. g., metals and hydrogen have a tendency to give up electrons and become positively charged ions; while non-metal (oxygen, chlorine, bromine, iodine) possess a tendency to take up electrons and become negatively charged ions.

Such tendencies to give up or take up electrons depend also on the particular resulting substances formed: many resulting substances may be changed back simply by reversing the direction of the electric change (i. e., reversing the direction of the current), and such substances actually exert tendencies which oppose the tendencies of the original substances.

Hence the force of a pole depends on the nature of the original substance and on the particular kind of resulting substance formed during its action. Furthermore, the force depends on the *concentrations* of the original and of the resulting substances: both the forward and the reverse tendencies to change increase somewhat with the amounts, per cc., of the original and of the resulting substances respectively. Hence, in recording measurements of the electromotive forces of poles, we must state the exact concentrations of the original and of the resulting substances in the poles when the measurements are made. Note how this is done in the table on page 110.

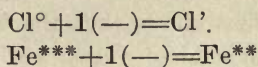
These pole-changes may be expressed by equations. Thus for copper and zinc, which are used in the demonstrations above, the equations are:



The Cu and Zn on the left side of the equation are marked with a *zero* to denote the fact that they are elements, i. e. they have

no ionic charges. An electron by itself is represented by (—), and the expression $-2(-)$ denotes that two electrons are given up by one atom of copper or of zinc.

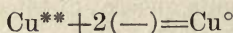
For the other two substances, the expressions of the pole changes are—



by which is meant that these substances *take up* one electron per atom and become *chlorine* ion and *ferrous* ion respectively.

But although a particular “pole” may have a *tendency to give up* electrons, it *cannot actually do so unless* it is connected with another pole which will *take up* electrons. When the “copper-copper sulphate” cup is placed in the connecting salt trough, together with the chlorine cup, and the poles are connected by a wire, then the copper actually changes to Cu^{**} ions because the electrons given up by the copper are “transferred” in a sense through the wire connection to the chlorine pole and are there taken up by neutral chlorine, which becomes Cl' ions.

In this particular combination both elements change to ions when the current flows. But when a “copper-copper sulphate” and a “zinc-zinc sulphate” pole are combined into a cell, then only one of these can change from metal to ions, i. e., only one metal can give up electrons, and the other takes up electrons and changes in the reverse manner, as represented by this expression—



The direction in which a pole actually changes when it is combined with another pole to form a cell depends upon the relation of the tendencies of the two poles. This relation has been ascertained for all poles by ascertaining the relations of all other poles to a common “reference” or “zero” pole: since the tendency of a single pole is independent of the other pole with which it is combined, the relation between this reference pole and all other poles gives also the mutual relations between these other poles. This procedure was illustrated in the demonstration above. All the data and further information needed in this connection is

Note.—An asterisk designates *positive* electric charges; an apostrophe, *negative* charges.

given in the "Table of Electromotive Forces of Battery Poles," page 109, which see.

Further general consideration of the galvanic cell need not be given here, but will be given in connection with the subject of oxidation and reduction (which see). It remains only to point out the connection between battery cells and electrolytic cells.

The substances produced by electrolysis (hydrogen from acids or water, copper from copper salts, chlorine from chlorides, etc.) naturally exert a tendency to regain the ionic state, thus each pole acts as the pole of a battery with an electromotive force opposite to the applied force. This is the electromotive force of polarization. It is absent before electrolysis begins and hence at the beginning any applied force, however small, will produce a current; but as soon as products of the electrode discharges are present, their electromotive force is exerted, and in order that electrolysis may continue, the applied force must be larger than this opposing force.

*The Action of General Reagents Upon Solutions of Salts.**

The action of reagents, the results of which depend only on our *general* knowledge of solubility and ionic dissociation, has already been illustrated above. Many other reagents, however, present *special* conditions of ionic dissociation which require separate consideration. Of the reagents which present special conditions, only those commonly used will be considered. It will be found that the *special* properties of the latter are also *essential* parts of chemical *information*.

1. *Sodium (or Potassium) Hydroxide as a Reagent.*

Experiment. Secure solutions of water soluble salts of the following cations: Cu, Ag, Zn, Cd, Hg (ous), Hg (ic), Pb, Fe (ous), Fe (ic), Ni, Al, Mg, Ca.

To a few cubic centimeters of each of these solutions in a test-tube add a few drops of sodium hydroxide solution, shake in order to mix thoroughly, observe the effect, add a few drops more of the reagent, and so continue until the reagent has been

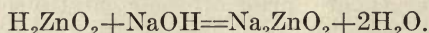
*Read pages 49 and 50 in this connection.

added in a relatively large amount, say about twice as much reagent by volume as of salt solution, if the two solutions are of approximately equivalent concentrations. Compare the results with the statements in the table below.

General Facts. By strict metathetical reaction the *hydroxides* should be obtained in the mixtures below. However, in many cases the substances finally obtained are not the hydroxides of the metals but substances derived from them through one or both of the two following changes:

(a) Dehydration,—complete ($\text{Hg}(\text{OH})_2$ to HgO) or partial ($\text{Cu}(\text{OH})_2$ to Cu_2O or $\text{Cu}_2(\text{OH})_2$).

(b) Dissolution of the precipitated hydroxide by excess of the reagent, thus showing that the precipitated hydroxide functions as an acid, e. g. $\text{Zn}(\text{OH})_2$ dissolves in excess of NaOH solution as per equation:



The formula for zinc hydroxide is thus written to *suggest its functioning as an acid*.

Table showing Results of Action of NaOH (or KOH) Solutions upon "Water Soluble" Salts of the Following Metals:

Ba—Ba $(\text{OH})_2$ is precipitated only from concentrated solutions because it is soluble in 20 parts of water—white.

Sr—Ppt. Sr $(\text{OH})_2$ sol. in 60 parts of H_2O —white—not precipitated from dilute solutions.

Ca—Ppt. Ca $(\text{OH})_2$ sol. in 700 parts H_2O —white—not precipitated from *very dilute* solutions.

Mg—Ppt. Mg $(\text{OH})_2$ sol. in 6000 parts H_2O —white.

Al—Ppt. Al $(\text{OH})_3$ white, sol. in excess of reagent, giving NaAlO_2 .

Zn—Ppt. Zn $(\text{OH})_2$ white, sol. in excess— Na_2ZnO_2 .

Pb—Ppt. Pb $(\text{OH})_2$ white, sol. in excess— Na_2PbO_2 .

Ferrous—Fe $(\text{OH})_2$ white ppt., darkens on exposure to air (oxidizes).

Ferric—Fe $(\text{OH})_3$ reddish brown, flocculent ppt.

Mn—Ppt. Mn $(\text{OH})_2$ —"flesh" colored, darkens on exposure to air.

Ni—Ppt. Ni $(\text{OH})_2$ —pale green.

Cu—in cold solution, $\text{Cu}(\text{OH})_2$ —bluish white ppt., soluble in large excess of reagent; in hot solution, CuO —black ppt.

Cd— $\text{Cd}(\text{OH})_2$, white ppt.

Bi— $\text{Bi}(\text{OH})_3$, white ppt.

Ag—Ppt. Ag_2O —greyish brown.

Hg(—ous)—Ppt. Hg_2O —black.

Hg (ic)—Ppt. HgO —yellow—in order to avoid the formation of different colored basic salts, the Hg salt solution must be poured *into* the reagent.

Exercise (a). In what way can it be readily ascertained whether or not a substance has been completely precipitated, say $\text{Fe}(\text{OH})_3$ by NaOH ? Hence would sodium hydroxide be employed as a precipitant for aluminium if it be required that precipitation be complete? From the metals given in the table above select those for which sodium hydroxide may thus be used as a precipitant.

(b) Since it is easy to convert a hydroxide completely to any salt by treatment with the necessary acid, and there is no troublesome (?) by-product formed, it is plain why the precipitation of metals in the form of hydroxides is very desirable. By means of such an intermediate step many metals may be readily changed from one salt to another—a change that may not be possible in one step otherwise. Give directions for changing the magnesium in magnesium sulphate to magnesium chloride; also for changing copper in copper nitrate to copper chloride.

(c) Given a solution containing aluminium and ferric salts. State how these metals may be obtained separately in the form of any compound. Similarly, how may zinc and cadmium be separated?

2. Ammonia as a Reagent.

Preliminary Considerations. Since a solution of ammonia contains OH ions we expect to find that it reacts just as sodium hydroxide does in the preceding exercise; that is, as a reagent to precipitate metal hydroxides. However, it differs from sodium hydroxide in being a weak base, and hence it has no effect

on solutions of salts of the strong bases (which are these?). With reference to the other bases, the following holds. *Ammonia precipitates the hydroxides of the trivalent metals completely under all conditions, but it does not precipitate these of the other weak-basic cations if enough of an ammonium salt (i. e., many NH_4^+ ions) is present* (exceptions—lead and mercury, which with ammonia always form insoluble, complex ammonium compounds). This influence of the NH_4^+ ion is exerted as follows: since ammonia is very slightly ionized, the product $(NH_4^+) \times (OH^-)$, which holds an equilibrium with the undissociated NH_4OH , must remain practically constant; hence when (NH_4^+) is increased by the addition of an ammonium salt (e. g. NH_4Cl) which is largely dissociated, (OH^-) must decrease, or in other words, *ammonia is less dissociated in the presence of an ammonium salt than in the absence of the latter*. The concentration of the OH^- ion in such a mixture is so small that the product obtained by multiplying this OH^- concentration with the concentration of any bivalent metal ion that may be added ($[M^{2+}] \times [OH^-]$) will be less than the "ion-product" that a resulting precipitate itself could produce. The formation of a precipitate does not take place because thereby the ion-product would be increased, and reactions take place only when an ion-product is decreased. But the ion-products of the tri-valent-metal-hydroxides are smaller than those of the bivalent-metal-hydroxides, and whenever a tri-valent-metal-ion is introduced into a solution containing any OH^- ions, the product $[M] \times [OH^-]$ is always greater than the ion-product of the precipitate,—hence the precipitate is formed because thereby the product of these ions in this solution is reduced.

To demonstrate this effect, add some ammonia to one portion of magnesium chloride solution, and to another portion add some ammonium chloride solution and then ammonia.

Exercise. By the use of ammonia, separate bismuth from a solution containing bismuth and copper salts; Al from a solution containing Al and Ni salts; Fe (ic) from a solution containing ferric and Mg Salts.

Give directions for changing the aluminium in potash alum completely to aluminium chloride.

3. Soluble Sulphides as Reagents.

Note. Hydrogen sulphide is poisonous, hence the generators of this gas should be used either outside of the building or in well drawing hoods.

Preliminary Considerations. Since all sulphides except those of the alkali and alkaline earth metals (Na, K, " NH_4 ," Ca, Sr, Ba, Mg) are insoluble, it is to be expected that a precipitate of the corresponding sulphide will be obtained whenever a soluble sulphide is added to a solution of a salt of any of the other metals. And such is the case except when hydrogen sulphide is used. Judging from its properties, this substance appears to be an exceedingly weak acid—very slightly dissociated—and its ion product $(\text{H}^*) \times (\text{S}'')$ is very small. A great increase of H^* ions, through the addition of a strong acid, lessens its S'' concentration greatly for the same reason that an increase of NH_4 ion concentration decreases the OH' concentration in ammonia. Hence the following general fact concerning the effect of hydrogen sulphide as a reagent:—

Even in the presence of a moderate concentration of hydrogen ions (a strong acid), hydrogen sulphide precipitates completely the sulphides of some metals, among them Ag, Hg, Pb, Bi, and Cd, which for convenience may be called the hydrogen sulphide group; and only in the absence of hydrogen ions (absence of acid) does it precipitate completely the sulphides of other metals, among which may be mentioned Fe, Zn, Ni, Mn, the ammonium sulphide group.

Demonstration. Put 10 to 20 cc. of dilute copper sulphate solution and an equal amount of zinc sulphate solution in a small flask, add 10 to 20 drops of dilute HCl , heat to boiling, treat with H_2S in excess and filter. What is the substance on the filter paper? The reaction that has taken place is metathetical—write the equation. What is the by-product? Evidently the concentration of the hydrogen ion has been increased during the progress of the reaction. Since with excessive concentration of hydrogen ion, hydrogen sulphide is unable to precipitate sulphides of this group even (?), and since it is possible that in this experiment the concentration of the acid has become excessive, the last portion of copper may not

be precipitable under the conditions. Hence a small part of the solution must be diluted largely by adding an equal volume or more of water, or the acid may be partially neutralized by the cautious addition of a base. Then it should be heated again and treated with hydrogen sulphide. If necessary, treat the main portion of the solution in this way and repeat this procedure until precipitation is complete.

Before proceeding to precipitate the zinc sulphide, it is desirable to remove the remnant of hydrogen sulphide which remains dissolved in the liquid, since otherwise it starts precipitation as soon as the acid is neutralized by ammonia or any other base. To remove this gas, boil the liquid well for two or three minutes.

Then add ammonia to the clear (and odorless) solution until it is plainly alkaline. If enough ammonium salt is present (from what source?), then no precipitate will be formed. Next, treat the solution with hydrogen sulphide again. A *white* precipitate (ZnS) will be formed (a trace of iron compounds introduced by the hydrogen sulphide may impart a "dirty" color to the precipitate). Filter and wash the precipitate.

The ammonia added above performs *two* functions: first, it neutralizes the free acid present, and second, it forms ammonium sulphide with the hydrogen sulphide. Write the metathetical reaction of the action of ammonium sulphide *with the zinc salt*.

Transfer the CuS to a dish, add a little dilute HNO_3 , and warm the mixture. A solution of $\text{Cu}(\text{NO}_3)_2$ will be obtained by complex reaction. Note the free sulphur formed.

Drip dilute HCl slowly upon the ZnS on its filter paper until all has been dissolved and collected in a beaker placed below. The reaction is metathetical(?). Note the evolution of H_2S .

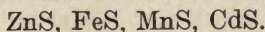
What is the object of this whole experiment?

Reactivities of Sulphides.

The reagents under (b) and (c) below will also react upon the sulphides in the groups above their own; but the reagents of (a) and (b) *do not* react upon the sulphides in the groups

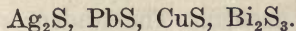
below them. In all cases the metals dissolve because they are changed to water soluble salts.

(a) Sulphides which react metathetically with dilute HCl or dilute H_2SO_4 :—

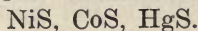


CdS requires a stronger "dilute" HCl than the others.

(b) Sulphides which react with dilute HNO_3 —action complex, resulting in free sulphur and nitrates of metals:—



(c) Sulphides which react with aqua regia only (aqua regia is essentially a concentrated solution of *free chlorine* which acts to produce free sulphur and chlorides of the metals):—



Colors of Sulphides.

The colors of sulphides are a valuable means of identifying the metals. By *proper means*(?) prepare small portions of the colored sulphides. They have the following colors:—

FeS, NiS, Ag_2S , HgS, PbS, Bi_2S_3 —dull black.

CuS—brownish black.

ZnS—White (when pure!).

CdS—yellow.

MnS—pink or flesh color.

Note:—during precipitation HgS frequently exhibits several other colors—black, yellow, red—but with excess of H_2S it finally becomes black.

Exercise. For how many metals may soluble sulphides be used as precipitation reagents? Into how many classes may metals be separated by means of soluble sulphides? The answers to these two questions will indicate why soluble sulphides are the most valuable general reagents known.

Precipitate separately as sulphides the metals from a solution containing any one of the following pairs of metals—copper and zinc, cadmium and nickel, bismuth and manganese. Continue the treatment for the precipitation of the first metal until the second sulphide can be obtained *pure*—as shown by its color. Wash the copper, cadmium, or bismuth sulphide with distilled water while it is on the filter paper (to remove remnants of the salt of

the second metal), and convert it to the nitrate. Convert the zinc, nickel, or manganese to the chloride.

List of Useful Special Properties and Reactions of Metals.

These properties and reactions are only briefly indicated here. Consult a text book for further information. These properties and reactions need not be definitely remembered, since they may be looked up when needed. The student should make simple trials to acquaint himself experimentally with these facts.

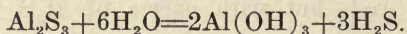
1. AgCl is soluble in ammonia, from which solution AgCl may be reobtained by neutralizing the ammonia with an acid (HNO_3).

2. PbCl_2 is very soluble in *hot* water, though not in cold water.

3. HgCl forms a black compound with ammonia, for which the formula and reaction need not be learned.

4. HgCl_2 is reduced to HgCl or Hg by SnCl_2 solution, the latter becoming SnCl_4 .

5. Bi salt solutions hydrolyze when the concentration of free acid (H^+ ion!) in the solution falls below a certain limit. "Hydrolysis" is a metathetical reaction between water and a salt which produces the free acid and the free base (or a *basic* salt), e. g. $\text{BiCl}_3 + \text{H}_2\text{O} = \text{BiOCl} + 2\text{HCl}$.



All salts of weak acids or of weak bases are hydrolyzed, though some only so slightly that the effect can not be noticed except by some delicate means such as the effect upon litmus. For demonstration, test solutions of the following *normal* salts with litmus:

Sodium carbonate.

Sodium acetate.

Copper sulphate.

Zinc chloride.

6. Note that the following salts and their solutions are colored:—

Copper salts—blue or green.

Nickel salts—blue or green.

Ferrous salts—pale green.

Ferric salts—reddish yellow.

Manganous salts—pale amethyst.

7. Compounds of the following metals will color the Bunsen flame. To try this use a *clean* platinum wire, dip it into solutions of these metals and hold the drop of the solution in the flame:—

Bright red—Sr.

Brick red—Ca.

Yellow—Na.

Yellowish green—Ba.

Green to blue—Cu.

Blue, pale—Pb.

Violet—K.

8. Ammonia in compounds is revealed by treating them with a strong base (e. g. NaOH solution), warming the mixture, and noting the odor.

9. Sodium hydroxide cannot be used as a precipitating reagent in the presence of ammonium salts because it reacts with the latter. Ammonium salts may be removed by evaporating the solution to dryness and then heating the dish and contents to low redness until fumes cease to be given off.

Chemical Problems.

The following problems are to be solved by means of the foregoing facts:

1. Students should be given solutions of water soluble salts, each solution should contain only one metal, and the student should ascertain what the metal is. Salts of the following metals may be given—copper, silver, bismuth, lead, mercury (both valences), cadmium, iron (ferric only) manganese, nickel, zinc, calcium, strontium, barium, magnesium, sodium, potassium, ammonium. Twelve to fifteen solutions should thus be worked out by every student.

In trying to find the metal, the student should note the color of the solution; he should ascertain if it would color the flame of the Bunsen burner; and on very small portions of the solution he should try the effect of reagents in the order given below.

(a) Add dilute hydrochloric acid. If a precipitate is obtained, then special tests 1 to 3 should be tried on the ppt. *after* it has been filtered off and washed.

(b) Add dilute sulphuric acid.

(c) Acidify moderately with either HCl or H_2SO_4 (which ever produces no precipitate), and then treat with hydrogen sulphide (see "g" below).

(d) Add ammonium chloride and ammonia. Ammonium chloride will produce a ppt. if HCl did, but such a ppt. should be removed by filtration before ammonia is added.

(e) Irrespective of the presence or absence of any ppt. produced by *ammonia*, treat the resulting mixture from (d) with hydrogen sulphide (see "g" below).

(f) Add sodium hydroxide.

(g) If a black sulphide has been obtained, then to decide which metal sulphide it is, try special tests 4 and 5 above, and also ascertain the reactivity of the sulphide with acids *after* it has been filtered off and washed.

2. Prepare KNO_3 by the commercial method, from NaNO_3 and KCl. (See *Smith*, Art. 127.)

3. Prepare NaOH by the commercial method, from Na_2CO_3 and Ca(OH)_2 . (See *Smith*, Art. 126.)

4. Prepare pure sodium chloride by precipitation with HCl. (See *Smith*, Art. 132.)

Text-Book Reading.

Parallel with this laboratory work on the metals some of the usual text-book reading found under the various metals should be given. However, it should be more or less confined to the *commercially important* facts and compounds, such as the commercial sources, and the methods of preparation, from these sources, of the commercially important products.

The choice of what is commercially important varies some-

what with the locality. In an agricultural state such as Texas the source of potassium salts and the composition of fertilizers is of vastly greater importance than the source of copper, and the methods of its extraction from its ores; while in a mining state such as Arizona just the reverse holds good. The emphasis in the informational reading should be laid accordingly.

Chemical Changes Involving Oxidation and Reduction.

Ionisation and Valence. In contradistinction to the changes studied thus far (metathetical changes and hydration) in which the *valence* of all ions (or constituent parts of compounds) *remained constant*, the changes now to be studied specially are those *involving changes of valences*. As has been shown, valence means the number of electric charges on an ion (or that would be on any particular part of an acid, base, or salt after it had been changed to an ion). Thus the valence of the NO_3 ion is 1 because in the ionization of any one of its compounds—e. g. NaNO_3 or HNO_3 , the *Na* or *H* gives one negative ionic charge (electron) to the NO_3 .

So far the attention has not been directed to the *complete* ionisation of such complex compounds as HNO_3 , NH_4Cl , CuSO_4 , etc., which would yield each element of a compound in the form of a separate ion, and this must now be taken up. Thus the primary ionisation of NH_4Cl yields NH_4^* and Cl' , but complete ionisation into the elemental parts requires the further ionisation of NH_4^* . Since 4H^* result from this ionization, and thus 3 new (+) charges appear, it follows that 3 (—) have been given up by the H's and left upon the N ion (written $\text{N}^{3'}$ or N'''). The whole compound in the form of ions may be expressed thus (4H^* , $\text{N}^{3'}$, Cl'). In the same way the complete ionisation of H_2SO_4 requires, after the primary ionisation into 2H^* and SO_4'' , the further ionisation of SO_4'' . Since $4\text{O}''$ would result from this secondary ionisation, and hence 6 *new* (—) charges appear, then the 6 (+) produced simultaneously

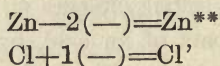
Note.—An asterisk designates positive electric charges; an apostrophe, negative charges.

must reside on the S ion (written S^{6*} or S^{*****}). The whole may be written $2H^*$, S^{6*} , $4O''$.

It is a general fact that in aqueous solutions hydrogen and metals become cations while O (or OH), the halogens (F, Cl, Br, I) and the cyanogen radical (CN) all become anions.

It has been shown above that in electrolytic and battery cells, during action, the substances at one pole undergo one kind of a valence change while the substances present at the other pole undergo the *opposite* kind of a valence change.

The proportion of the substances changing valences simultaneously at the two poles depends merely upon the fact that the number of electrons simultaneously transferred "at" the two poles *are equal* (but they are transferred in opposite senses at the two poles). Thus if the two "pole changes" are



then the proportion in which the two substances change simultaneously is $Zn:2Cl$.

When the results produced upon a substance by these changes in valence are compared with the results *produced upon it* by oxidizing (or reducing) agents, it is found that an increase of positive ionic charges (or decrease of negative) produces the same result as *oxidation*, and the reverse electric change produces the same result as *reduction*. This has been found to be always true. For illustration, consider the production of chlorine from HCl. Hence the following simple and perfectly general definition:—oxidation is the increase of positive charges or valences of an element or radical (or the decrease of negative charges). Reduction is the reverse change.

Exercise.

Ascertain whether the elements *italicized* when changed from the first to the second form undergo oxidation or reduction. Express the amount of change in numbers of ionic charges per atom of the changed element.

First Form.

- (a) KCl
- (b) *Chlorine*
- (c) $ZnSO_4$
- (d) *Zinc*
- (e) K_2SO_4
- (f) HNO_3
- (g) $KMnO_4$
- (h) $FeSO_4$
- (i) K_2CrO_4

Second Form.

- HCl
- HCl
- $ZnCl_2$
- $ZnCl_2$
- KCl
- NH_4OH
- $MnSO_4$
- $FeCl_3$
- $CrCl_3$

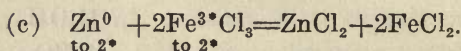
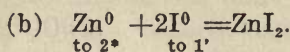
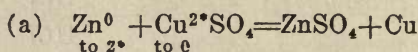
In the light of this definition of oxidation and reduction, it is seen that in *every* electrolytic cell and every battery cell, oxidation takes place at one pole, and reduction at the other.

Study next the "Table of Electromotive Forces of Battery Poles," page 109.

If the "valence-changing" substances of the two poles of a battery cell *are put in direct contact*, (i e. mixed together) they will undergo the same changes *without the use of the metallic poles and their connecting wires*. Here evidently the transfer of the electrons is direct from one valence-changing substance to the other. This is easily shown by the following test-tube trials with some of the same substances used in the battery cells shown before:—

- (a) Put granulated zinc into some copper sulphate solution, and shake the mixture at intervals until the solution is colorless. Then test solution with H_2S for zinc ion. Conclusion?
- (b) In place of chlorine in a chloride solution it is *better to* use iodine dissolved in KI solution (or bromine water). Treat this with zinc until the free iodine (or bromine) has disappeared. Then test the solution for zinc ion. Conclusion?
- (c) Treat a ferric chloride solution with zinc until the reddish brown color of the ferric ion has disappeared. Test with NH_4Cl and NH_4OH to ascertain if the change is complete.

The three equations for these reactions are:—



The numbers of electric charges on the *valence changing* substances are written as shown in order to show at a glance which substances are oxidized and which are reduced. Point out which are oxidized,—which reduced.

The proportion between the valence-changing substances in a reaction is determined by making the number of electrons given up by one substance equal to the number of electrons taken up by the other substance. Thus in (b) above, one Zn gives up 2(—) while one I takes up only 1(—); hence their ratio in the reaction is 1Zn to 2I. Another way of stating the fact is to say that the amount of oxidation in a reaction is equal to the amount of reduction.

Exercise:

Make test tube trials of the following substances:—

1. Treat ferric chloride solution with excess of hydrogen sulphide, and test with ammonium hydroxide.
2. Add a slight excess of bromine water to a ferrous salt solution and test with NH_4OH .
3. Treat dilute bromine water with hydrogen sulphide.
4. Put a rod of zinc into a concentrated solution of stannous chloride.
5. Put a piece of copper into some mercuric chloride or nitrate solution.

Write the equations for these reactions as in the illustration above.

The foregoing reactions are evidently quite simple, yet *they present all the principles involved*, and other reactions which *appear to be* more complex really involve no new facts. As a

rule they differ only by involving some additional metathetical actions. This is well illustrated by the oxidizing actions of HNO_3 , which will now be discussed.

Nitric acid and its Reduction Products. The following should be among the experiments shown: —

1. The preparation of HNO_3 by distillation from NaNO_3 and concentrated H_2SO_4 .
2. The ease with which HNO_3 gives up oxygen is *strikingly illustrated* by dropping a glowing piece of charcoal into some hot *fuming nitric acid* in a test tube.
3. The preparation of NO by the reduction of dilute HNO_3 with copper. To be performed as usual. Show that NO combines with oxygen to form NO_2 and that this dissolves in *cold* water forming $\text{HNO}_3 + \text{HNO}_2$.
4. The *extreme reduction* of the nitrogen in the nitrate ion by means of a metal very high in the electromotive force table (see page 110),—e. g. zinc or aluminium: take concentrated sodium hydroxide solution, add a little of a nitrate, then add either one of these metals in the form of shavings or powder. Warm the mixture. After violent reaction has set in, the odor of ammonia can be easily noticed.
5. Treat some cold, freshly prepared, ferrous sulphate solution with dilute HNO_3 . A black compound (of FeSO_4 , NO) will be formed. Heat the mixture to boiling. The NO escapes, and the reddish brown color of the solution shows the presence of ferric iron.
6. Pass H_2S through warm dilute HNO_3 . Sulphur and NO will appear.

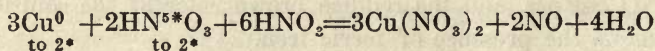
In order to consider properly the oxidizing actions of HNO_3 , the following facts must be noted:—

The nitrogen in nitric acid and in nitrates is in the *highest* state of oxidation in which it ever occurs; and the nitrogen in ammonia or ammonium compounds is in the *lowest* state of oxidation,—it cannot be reduced further. Between these two limits there are a number of oxidation stages, the order and relation of which is shown in the following table. Note that the free element occupies an intermediate position.

Name.	Nitrogen Ion.
Nitric acid, HNO_3 and nitrates.....	N^{5*}
Nitric peroxide or nitrogen tetroxide, NO_2 or N_2O_4	N^{4*}
Nitrous acid, HNO_2 and nitrites.....	N^{3*}
Nitric acid or nitrogen dioxide, NO	N^{2*}
Hyponitrous acid, NHO , and nitrous oxide N_2O	N^*
Nitrogen, N_2	N^0
Ammonia, and its salts, $(\text{NH}_4)^*$	$\text{N}^{3'}$

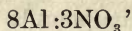
The equations for the experiments above may now be derived.

Experiment (3). When completely ionized, HNO_3 is H^* , N^{5*} , $3\text{O}''$). Since the nitrogen finally appears as NO ,—i. e. (N^{2*} , O'')—it has received 3(—), three electrons. Since these electrons are obtained from Cu^0 becoming Cu^{2*} , it follows that these two substances must change in the ratio 3Cu to 2HNO_3 in order that the number of electrons given up by the Cu shall be completely taken up by the changing HNO_3 . But when 2HNO_3 change (to 2NO), then 2H^* and $4\text{O}''$ will become free ions, and since O'' ions cannot remain free in the presence of H^* ions, they will form $4\text{H}_2\text{O}$. This requires 8H^* , for which the 2HNO_3 just considered supply only 2H^* , and 6 extra HNO_3 must give up their H^* ions. The $6\text{NO}_3'$ from these 6HNO_3 balance up with the 3Cu^{2*} formed in this action. Considering all this we write—



Which substance is oxidized? which reduced? Is the amount of oxidation equal to the amount of reduction?

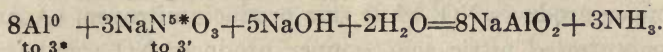
Experiment (4). In the change from nitrate ion to ammonia, each "N" receives 8 electrons, while each "Al" gives up 3 electrons; hence the ratio between the number of Al and the number of NO_3' which change simultaneously is



In the presence of sodium hydroxide, the Al^{***} ion combines with 1Na^* and $2\text{O}''$ to form NaAlO_2 . The $\text{N}^{3'}$ ion forms NH_3 ,

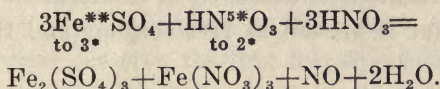
Note.—All aqueous solutions contain H^* and O'' ions, and when these free ions are used up more will be formed by further ionisation.

hence it combines with 3H^* . Considering all this, we arrive at the following equation:—

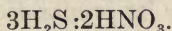


Which element or substance is oxidized? which reduced? Is the amount of oxidation equal to the amount of reduction? These questions should be asked in connection with all equations of oxidation and reduction.

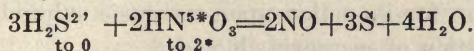
Exp. (5). One Fe^{**} gives up one electron to become Fe^{***} , and since the nitrogen is reduced from HNO_3 to NO (i. e., 3 electrons are taken up by 1 N), then the ratio between the valence-changing substances is $3\text{FeSO}_4:\text{HNO}_3$. The other considerations are the same as in (3) above. The equation arrived at is:—



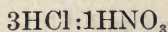
Exp. (6). S'' ions change to S^0 . Since NO is formed from NO_3' , the ratio between the valence-changing substances is



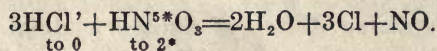
Hence the equation is:—



One more equation should be added here: the reaction of "aqua regia." Considering that NO is the reduction product and free chlorine is the oxidation product, the ratio between the valence changing substances must be



and the equation is



The foregoing arithmetic considerations are second in importance to the real chemical knowledge to be gained from the experiments above. The latter consists of knowing what particular products of oxidation or reduction are obtained under particular circumstances. This must be remembered. The "chemical" aspect of the experiments with nitric acid involves the following general facts (see also the Table of Electromotive Forces of Battery Poles):—

The *extent of reduction* of nitrogen in compounds depends

upon the strength of the reducing agent and upon the dilution of the nitrogen compound. For example, with dilute nitric acid, copper produces NO, while zinc produces nitrogen. Again, with *concentrated* nitric acid, copper produces NO_2 ; and with *very dilute* nitric acid, zinc reduces the nitrogen from nitric acid to ammonia.

These general relations should be impressed by drill with special examples as illustrated in the following exercise.

Exercise.

1. Figure out the equation of the reaction that takes place when copper sulphide is dissolved in dilute nitric acid if at the end the sulphur is present as free sulphur and the nitrogen as nitric oxide.

2. Do the same with bismuth sulphide, and again with silver sulphide.

3. If aqua regia, after its preparation by mixing the acids, is essentially a solution of chlorine (see equation above), what will be the equation for its action upon mercuric sulphide which results in a solution of mercuric chloride?

4. Note the position of silver, relative to copper, in the Electromotive Force Table, page 110, and hence infer what will be the probable reduction product of nitric acid if the concentrated acid reacts with silver. Write the equation for this action.

5. Do the same for lead and dilute nitric acid; and again for cadmium and very dilute nitric acid.

Note on the Oxidizing Action of Sulphuric Acid. Sulphuric acid is occasionally used as an oxidizing agent. The student may have occasion to meet it in connection with the dissolution of metals in it, or in its action upon potassium bromide and potassium iodide. Since the reduction products of sulphuric acid are all known to the student, it is advisable to add here the few remarks necessary to inform him fully concerning these particular reactions.

The order of the reduction products of sulphuric acid and of their ions is as follows:

Order	Substance	Ion.
1 Highest state of oxidation	H_2SO_4	$\text{S}^{\bullet\bullet}$
2. Next lower	H_2SO_3 or SO_2	$\text{S}^{\bullet}$
3. Next lower	S	S°
4. Lowest	H_2S	S''

With weak reducing agents such as metallic silver and metallic copper concentrated sulphuric acid will be reduced to SO_2 only. Of course the sulphate of the metal is produced in the reaction.

With the strong reducing agents such as zinc, the reduction product is H_2S .

Concentrated sulphuric acid does not oxidize chlorine from HCl ; but it oxidizes bromine from *bromides* forming free bromine, and the sulphuric acid itself is reduced the least amount only, namely to SO_2 . Iodine from *iodides* is so easily oxidized that the sulphuric acid is reduced to H_2S even.

For a numerical exercise, the first five equations for the five reactions just mentioned may be figured out.

Table of Electromotive Forces of Battery Poles.

In the table below, of electromotive forces of some battery poles, the voltages given are those of the cells formed by combining each of these poles with a hydrogen pole, which latter serves as a zero or reference pole.

The prefixed algebraic signs indicate the polarity of the "measured" pole in this combination.

The accompanying scale of these forces shows their mutual relation.

When a pole-component changes so as to be deprived of or give up electrons, then it is oxidized, or acts as the negative pole of a battery cell.

Any pole in the table combined with any other pole will form a battery cell the force of which is shown in the "Scale showing the relation of EMF's": it is represented by the distance between the poles. The pole uppermost in the table acts as the negative pole in this combination,—here oxidation takes place.

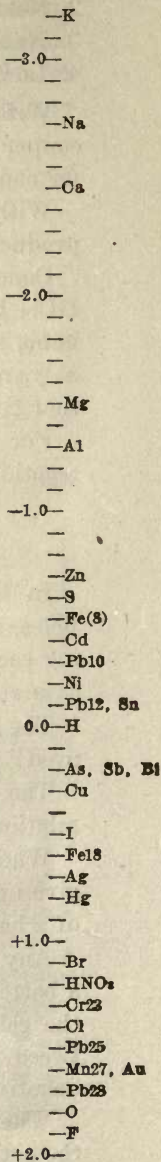
When the substance that would be used up at one pole during the action of a cell is mixed or placed in direct contact with the substance that would be used up at the other pole, then

Table of Electromotive Forces of Some Battery Poles.

No.	E. M. F.	Higher state of oxidation.	Direction in which change may take place.	Lower state of oxidation.
1	-3.2	N/1 K* solution	In both directions	Solid potassium, K°.
2	-2.8	N/1 Na* solution	In both directions	Solid sodium, Na°.
3	-2.5	N/1 Ca* solution	In both directions	Solid calcium, Ca°.
4	-1.6	N/1 Mg* solution	In both directions	Solid magnesium, Mg°.
5	-1.3	N/1 Al*** solution	In both directions	Solid aluminum, Al°.
6	-0.76	N/1 Zn** solution	In both directions	Solid zinc, Zn°.
7	-0.55	Sulphur, solid	In both directions	N/1 S** solution.
8	-0.43	N/1 Fe** solution	In both directions	Solid iron, Fe°.
9	-0.40	N/1 Cd** solution	In both directions	Solid cadmium, Cd°.
10	-0.34	N/1 PbSO ₄ solid saturating a solution of dil. H ₂ SO ₄ , sp. gr. 1.20.	In both directions	Solid lead, Pb°.
11	-0.17	N/1 Ni solution	In both directions	Solid nickel, Ni°.
12	-0.12	N/1 Pb** solution	In both directions	Solid lead, Pb°.
13	-0.10	N/1 Sn** solution	In both directions	Solid tin, Sn°.
14	+0.00	N/1 H* solution	In both directions	Hydrogen gas, H°.
15	+0.2(?)	Solution of As*, Sb*, or Bi**	In both directions	Solid As°, Sb°, or Bi°.
16	+0.34	N/1 Cu* solution	In both directions	Solid copper, Cu°.
17	+0.54	Iodine, saturated solution in N/1 iodide solution	In both directions	N/1 I* solution.
18	+0.75	N/1 Fe*** solution	In both directions	N/1 Fe* solution.
19	+0.80	N/1 Ag* solution	In both directions	Solid silver, Ag°.
20	+0.86	N/1 Hg** solution	In both directions	Mercury metal, Hg°.
21	+1.08	Bromine, saturated solution in N/1 bromide solution	In both directions	N/1 Br* solution.
22	+1.20(?)	Conc. HNO ₃ (N°*)	Left to right only	Different reduction products, according to conditions.
23	+1.3(?)	Acidified bichromate solution (Cr**)	Left to right only	Solution of chromic salt, Cr***.
24	+1.35	Chlorine, sat. solution in N/1 chloride sol. (see note below)	In both directions	N/1 Cl* solution.
25	+1.44	PbO ₂ , solid, (Pb****)	In both directions	N/1 Pb** solution.
26	+1.50	N/1 Au* solution	In both directions	Solid gold, Au°.
27	+1.50(?)	Acidified permanganate solution, Mn7+	Left to right only	Solution of manganous salt, Mn**.
28	+1.66	Solid PbO ₂ , (Pb****)	In both directions	Solid PbSO ₄ saturating a solution of dilute sulphuric acid, sp. gr. 1.20.
29	+1.7	Oxygen liberated on platinum	Right to left only	From neutral or slightly acid solutions of sulphates, nitrates, phosphates, etc.
30	+1.9	Fluorine gas, F*	In both directions	N/1 F* solution.

Note.—An asterisk denotes a positive electric charge; an apostrophe, a negative charge.

Scale showing relation of EMF's.



these two substances react in the same way as in the battery cell. Direct contact makes the connecting wires, etc., for the transfer of the electrons unnecessary. All mixtures undergoing oxidation and reduction reactions may be considered to be made up in this way.

The substances in the higher states of oxidation (in the wide column on the left) are all oxidizing agents: they appear in the *ascending order of strength as oxidizers*.

The reducing agents are in the column on the right: they appear in the *descending order of strength as reducers*.

Any reducing agent in the table will react with any oxidizer *below it* in the table. The following two common examples of this fact should be noticed:

(a) Any metal in the table will displace (reduce) any metal below it from a solution of its simple salts. Thus, zinc or iron will displace copper from copper sulphate, silver from silver nitrate, or hydrogen from hydrochloric acid.

(b) Anyone of the non-metals (Cl, Br, I, S), will be displaced (oxidized) from its simple-ion compounds (chlorides, bromides, iodides, sulphides), by any one of the non-metals below it in the table. Thus, chlorine liberates bromine from bromides, sulphur from sulphides, etc.

The *metals above zinc* in the table below, and *fluorine*, cannot be used as poles for practical batteries, or as reducing (respectively, oxidizing) agents in ordinary chemical (aqueous) mixtures because these elements react with water.

Note on the potential of chlorine: its potential is less noble ("noble" means the direction in the table from hydrogen towards gold) if the concentration of the free chlorine is less or if the concentration of the chloride compound is greater than as given in the table. Hence the formation of the first portions of free chlorine in a concentrated solution of hydrochloric acid will take place with a potential between +1.35 and +1.0 volts. This explains why nitric acid oxidizes hydrochloric acid.

Note on the "Back" Electromotive Force of the Poles During Electrolysis.

The products formed at the poles during electrolysis exert an "opposing" electromotive force, i. e. "opposed" to the ap-

plied force. They exert about the same force as the battery poles composed of the same materials.

Of different chemical changes possible at the cathode of any particular cell that one will take place which will exert the least "zincic" potential (zincic means the direction in the table from hydrogen toward zinc).

At the anode that particular change will take place which requires the least "noble" potential.

For the liberation of hydrogen on metals other than platinum, a more zincic voltage is required than for platinum. This additional voltage amounts to as much as 0.5—0.7 volts for the metals lead, zinc, and mercury.

APPENDIX.

*The Details of Construction, Action and Operation of Alternating Current Rectifiers.**The Action of the Electrolytic Cell.*

The electrolytic cell used to "rectify" the alternating current, as ordinarily constructed, has one pole of aluminium, the other of lead, and between them a solution of sodium phosphate. When prepared for operation, the cell allows only that alternation of the current to pass which makes the aluminium pole the cathode, or negative pole, and it prevents the passage of the reverse alternation, which would make the aluminium the anode or positive pole. This property of the cell is due to the fact that the aluminium is covered with a non-conducting film of aluminium hydroxide or basic salt with a gas (oxygen) held in its "meshes" or pores. Such a film over a metal does not prevent the passage of the current when this metal serves as the cathode; but it hinders the passage of the current when this metal serves as an anode. It is as though the cations need not actually touch the metal of the pole in order to be discharged, while the anions cannot be discharged unless they come in direct contact with the metal surface. Hence a cell in which one pole is covered with such an insulating film and the other not covered, or with a "conducting" surface, is "uni-directional" in its action.

The materials used for this purpose, aluminium and lead, are particularly well suited for this purpose on account of the following advantageous properties possessed by them:

(a) when a bare surface of aluminium is exposed to anodic discharge (in solutions of phosphates, sulphates, borates, etc.) a film (of aluminium hydroxide) is almost immediately formed; and this film is not disturbed by any cathodic discharge such as takes place in sodium salt solutions. Hence when a bare aluminium pole is connected with an alternating current, that alternation which makes it the anode will quickly cover it with such an insulating film, and the formation of this

film will not be hindered by the reverse alternation of the current.

(b) a lead pole, which in the rectifier acts mostly as the anode, has its surface converted to lead peroxide by the anodic discharge, and this material, in contradistinction to the aluminium film is a good conductor. If the lead acts as a cathode, this peroxide is reduced to metallic lead. Thus neither anodic nor cathodic action covers the lead with an insulating layer.

With high voltages, or hot solutions, the aluminium film breaks down frequently in its weaker spots, and considerable leakage of current occurs. Free acid in the electrolyte also increases the leakage current. But with cold, neutral solutions and a voltage below 70 volts, the leakage current is not excessive.

Construction of Large Electrolytic Jars for Rectifier Set No. 1.

Secure 2 jars of 2 to 4 quarts capacity, with a depth of from 7 to 11 inches. Have a carpenter cut 4 discs of one-inch pine lumber, two discs of a size to fit inside of the upper edge of the jar, and the other two to fit over the top of the jar. By means of screws fasten one of the smaller and one of the larger discs together with the grains of the two pieces at right angles to each other; cut a *rectangular hole* through the center of these discs of such a size that the aluminium plate which is to be used will just pass through this hole. One side of the hole should be cut practically vertical, the other should be cut slantingly so that a suitable thin wedge inserted on the slanting side of the hole may serve to hold the aluminium plate rigidly in a vertical position when this wooden cover is placed on top of the jar. Secure, by mail, from the U. S. Aluminium Co. (Pittsburgh, Pa, or Old Colony Bldg., Chicago, Ill.), specifying the softest grade for which their stock number is 2SO, four aluminium plates 3 inches wide, 12 inches long, 1-8 inch thick, 13-5 lbs., which will cost 80 cents. Drill a hole near one end of each of two plates to screw on a binding post. (It is almost impossible to fasten the connecting wires on the aluminium by soldering.) Another way to fasten the connecting wire is to drill a hole about the size of the wire through the

upper edge of the plate, put in the end of the wire, and clamp it by hammering the plate together at that point. Secure from a plumber four strips of sheet lead about 3 to 4 inches wide, 6 to 7 inches long and thick enough to be stiff. Flatten each out thoroughly, and at one end bend a strip $\frac{1}{2}$ inch wide, sharply at right angles to the rest of the plate; on this strip solder the end of a 6 to 10 inch piece of copper wire (No. 14). By means of brass screws driven through this strip into the wooden cover, fasten two lead plates to the under side of each cover, so that they will be parallel to the aluminium plate at a distance of $\frac{1}{2}$ to $\frac{3}{4}$ inch, one on each side. Then dip the wooden cover into hot paraffin, so that all the wood and the lead plate fastenings will be covered with it. Insert the aluminium plates, fill the jars three-fourths full with the phosphate solution (prepared as directed below), and make the connections shown in Fig. 6.

The two jars may be reduced to one by placing both aluminium poles in the same jar with one lead plate between them and one on the outside of each aluminium plate. Thus three lead plates are used, which are all connected to the same terminal. The connections are the same as with two jars.

Design and Action of Special Transformer.

The transformer to be used in this connection is made of only one coil. One connection for the current to be drawn from it is at the center of the coil (see Figure 6). The "positive" direct current flows outward from this, or in other words, this is the direct current + pole connection. In addition there are other pairs of connections to the coil, the two connections of each pair being "balanced" with reference to the center, i. e., between the center and one connection there are as many "turns of wire" as between the center and other connection of a pair. One connection of each pair may be connected with the aluminium pole of one cell, and the other to the aluminium pole of the other cell. The lead poles of the two cells are connected to a common wire from which the negative direct current flows, in other words, it is the direct current—pole connection.

The voltage between any two points of the primary coil of a transformer is practically such a fraction of the voltage of the primary circuit as the ratio that the number of "turns of wire" intercepted on the coil by the two points bears to the number of turns in the whole coil. Hence between the center and either extremity one-half of the primary voltage is obtained, and between the center and any other points, lesser voltages proportionate to the number of turns between it and the center are obtained.

The direction of the electric pressure (or sign of the voltage) on one side of the centre is the opposite of that exerted simultaneously at the other side. If at any particular moment the voltage on the left of the centre is such as to make the left connection negative, then this would tend to send a current into the (left) cell with which it is connected, thus making the aluminium pole of this cell the cathode, and the lead pole the anode. Under these conditions the aluminium pole opposes no hindrance to the passage of the current. However, exactly the reverse condition exists simultaneously at the connection to the right of the center of the coil and the aluminium pole in this second cell opposes (practically prevents) the passage of the current from this side. With the next alternation of the current all the conditions are reversed. This time the current is free to flow through the cell on the right, but it cannot flow through the cell on the left. Thus the current which flows out of the coil at the centre is always positive, but it comes from the right side of the coil at one moment, and from the left side at the next moment.

In ordering such a coil suitable for "Rectifying Sets No. 1" above, the following should be specified:—the transformer is to be of the single coil or "auto" type, and is to be designed for a "110 volts, 60 cycle" circuit; capacity, 5 amperes at 110 volts. It is to be without case, but equipped with frame castings in order that it may be fastened to the table or wall. *The following connections are to be made*—nine leads to be connected as follows:

Two, at the ends of the coil—for the main line.

One, at the centre.

Three pairs for connections to be "balanced" with reference to the central tap and connected to the coil so as to secure the following voltages (approximately):

13¾ volts between the center and either one of the first pair.

27½ volts between the center and either one of the second pair.

41¼ volts between the center and either one of the third pair.

In other words, the transformer is designed for eight circuits of 13¾ volts each.

The Moloney Electric Co. of St. Louis, Mo., will furnish such a transformer for \$13.00 or slightly less, the general appearance of which is shown on page 43 of their catalogue.

Preparation of Solution for Electrolytic Jars.

For each quart of solution use one-half pound of sodium phosphate crystals, and add phosphoric acid in small portions until the mixture does not turn litmus paper a *deep blue*, but a lighter shade—not pink, however.

Operation of the Rectifier Set No. 1.

When the current is to be turned on for the first time after the rectifier has been assembled (or for the first time after a long period of rest) the *lowest* alternating current voltage should be connected to the electrolytic cells, and this should remain connected at least several minutes before any higher voltage may be connected. After that no further precaution need be taken. In the absence of an insulating layer over the surface of the aluminium the current will pass direct through *both* cells, but the liquid resistance of the cells will be large enough to prevent this current from being excessively large *if the voltage is small*. This current in passing *forms* the films over the surface of the aluminium plates.

During action aluminium phosphate is slowly formed and appears as a white flocculent sediment. Thus the aluminium pole and the phosphate ion (the PO_4 radical) are used up very slowly. Hydrogen also is discharged at the cathode. To re-

place this loss, small amounts of phosphoric acid must be added to the solution from time to time, with the aid of litmus as in preparing the solution. Naturally, any water lost by evaporation must be replaced. Under these conditions the cells will operate practically indefinitely.

Depending upon the rate at which the current is drawn, the voltage of the direct current obtained will be from 10 to 25 volts less than the alternating current voltage applied to the electrolytic jars.

When a current greater than 5 amperes is drawn continuously for more than an hour, the liquid in the electrolytic jars is heated so rapidly that it may finally become too hot for safe or economical operation. Hence when large currents are needed for longer periods of time, the electrolytic jars must be cooled in some way. This can be done by replacing the glass jars by heavy tin or sheet iron cans of the same size (which can be made by any tinner), and placing these electrolytic vessels in a large vessel filled with cold water. In extreme cases, ice water may be used.

A Small Cheap Rectifier.

Secure a 1 quart fruit-jar or a similar vessel. Secure a strip of sheet lead about 3 inches wide and 10 to 12 inches long. Hang this lengthwise into the jar, along the wall, and clamp it in position by bending the upper end over the rim, and solder a connecting wire to this. Fasten a wooden cover (6x6 inches) on the top of the jar by means of pieces of wood nailed as cleats to the under side of the cover. Secure by mail, from the U. S. Aluminium Co. (Pittsburgh, Pa., or Old Colony Bldg., Chicago) rods of aluminium $\frac{1}{2}$ to $\frac{3}{4}$ inches in diameter, and 10 to 12 inches long. Specify the softest grade (Cat. No. 280). Four rods $\frac{1}{2}$ x12 inches will cost 50 cents.

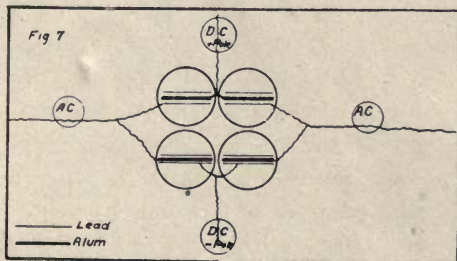
To fasten a copper connecting wire to the aluminium rods, drill, near one end of each, a hole large enough to admit the wire. When making connection insert the wire and clamp it by a few blows of a hammer on the side of the rod. Near the center of the cover, bore one (or two) holes of the exact size of the aluminium rods, so that the rods extending

through these holes will be held in vertical position. Connect this cell *in series* with either a lamp-bank or a suitable rheostat. The lamp-bank should have 8 to 12 incandescent lights of 32 candle-power, all connected *in parallel*. In place of this lamp-bank it will be more convenient to use a "theatre dimmer" for 12 lights (Ward-Leonard Electric Co., Bronxville, N. Y. (Cat. No. SD 16), price about \$2.50). After filling the jar with sodium phosphate solution, the rectifier is ready for use and can be used without any particular precautions.

How to Rectify Both Alternations Without a Transformer.

Without the aid of the special transformer, the only way to rectify both alterations is to use four simple cells and connect them as shown in Fig. 7. It is absolutely necessary to insert a large resistance in the alternating current circuit until the surfaces of the aluminium plates are formed. For this purpose cells of the size first described above require two 32 c. p. lamps. As the surface "forms" the lamps become dim. Then the lamps may be taken out of the circuit because the film is able to stand 110 volts. A single-pole, double-throw switch is usually used to facilitate this change. The direct current obtained from this rectifier will have a voltage of 80 to 95 volts. If, as is usually the case in laboratories, the current is to be used at a low voltage, a *lamp-bank* or some "high resistance" rheostat must be inserted in the circuit, either in the alternating or in the direct current side.

This form of rectifier, together with its necessary accessories, costs fully as much as the first rectifier described, and is less convenient and less suitable.



A Four-jar Rectifier (without transformer)

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